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EXPERIMENTAL CHEMISTRY.



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EXPERIMENTAL CHEMISTRY,

FOUNDED ON THE WORK OF

DR. JULIUS ADOLPH STÖCKHARDT.

A

HANDBOOK FOR THE STUDY OF THE SCIENCE
BY SIMPLE EXPERIMENTS.

By C. W. HEATON, F.I.C., F.C.S.,

LECTURER ON CHEMISTRY IN THE MEDICAL SCHOOL OF CHARING CROSS HOSPITAL,
EXAMINER IN CHEMISTRY TO THE ROYAL COLLEGE OF PHYSICIANS
AND THE ROYAL COLLEGE OF VETERINARY SURGEONS,
TREASURER TO THE SOCIETY OF PUBLIC ANALYSTS, ETC.

NEW EDITION, REVISED.

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PREFACE.

“STÖCKHARDT’S Principles of Chemistry,” in its English dress, has for many years filled a definite and useful place among elementary text books. It has appealed to the numerous class of students, both men and boys, who in spite of limited means and opportunities are anxious to acquire some experimental knowledge of the science—who intend to *work* at chemistry instead of merely *reading* about it. To such students it is useless to describe experiments which can only be performed with the aid of costly and elaborate apparatus, or with the skill derived from long practice.

The great merit of Dr. Stöckhardt’s book is that, while the experiments are clearly described and very numerous, they do not require for their successful performance any but the simplest and cheapest forms of apparatus. It is astonishing what an amount of good work may be done in chemistry with Florence flasks, tumblers, medicine bottles, basins, saucepans, tin plate, iron wire, corks and other articles, which are always at hand. If to these are added some glass and caoutchouc tubing, a few funnels, test-tubes and beakers, a mortar and pestle, spirit lamp (if gas cannot be used), a small set of scales and weights,

and a measuring glass, the student will be able to perform the great majority of the elementary experiments of chemistry.

Of course other apparatus may be added with great advantage, and apparatus is so cheap in the present day that almost every student will find it possible to make such additions to his stock as he goes on. But it is better not to begin with too much, as the beginner is almost sure to order many things which he will not really require.

When the publishers put Dr. Stöckhardt's work in my hands, I hoped that it might be possible to bring it into accordance with modern chemical ideas, without entirely altering its plan. This soon proved to be impracticable, and I found myself compelled to re-cast, and to a great extent to re-write the book. In doing so I have tried to preserve as far as I could the spirit of the original; but so much has been added, so much subtracted and so much altered, that I have not thought it right to retain the original title of the work. Nearly the whole of the original work which remains consists of experimental details and technical descriptions, which, though they have required much pruning and revision, are as useful now as when first written. A great many new experiments have been added in every part of the work.

The systematic classification adopted in Parts II. and III. has not been attempted in Part IV. (Organic Chemistry), because the more old-fashioned arrangement seemed better adapted for the purposes of a simple experimental study. But I have endeavoured, in two introductory chapters, to give the student some idea of the modern

system of classification, and have, in the subsequent pages, referred constantly to these chapters.

In Part I., a somewhat lengthy chapter (Chapter III.) has been devoted to the laws of chemical force. This chapter is longer than it might otherwise have been, because I have thought it right to establish the laws of constitution and chemical action, and the primary meaning of symbols and formulæ, *before* taking advantage of the great assistance of the atomic hypothesis. I trust I have succeeded in proving to the careful student that the symbols and formulæ of chemistry might be used with propriety and advantage, even if the atomic hypothesis were swept away.

In Parts II. and III. I was much assisted by Mr. E. Francis, F.C.S., Demonstrator of Chemistry in the Medical School of this Hospital. It is a pleasure to me to take this opportunity of thanking him for his valuable help in this and in many other things.

C. W. H.

CHARING CROSS HOSPITAL,
Christmas, 1871.

PREFACE TO NEW EDITION.

AFTER careful consideration I have decided to make no change in the plan of the book, in the part devoted to Organic Chemistry. Of course many alterations in detail have been necessary, but I still think that for students working with limited means, the system adopted by Dr. Stöckhardt, although now old-fashioned, is, on the whole, the best.

C. W. H.

CHARING CROSS HOSPITAL,
July, 1886.

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SIMPLIFIED FROM THE TABLE OF MR. FRANK WIGGLESWORTH
CLARKE.

THE atomic weight of an element, denoted by its symbol, is the number which most nearly fulfils the following conditions:—

1. It is the least weight of the element ever found in two volumes of any gas, elementary or compound, at standard pressure and temperature; the least weight of hydrogen so found being taken as unity. When greater weights are found they are always multiples of the least weights.

2. It is the weight of the element that is equivalent, in combination or substitution, to 1, 2, 3, 4, &c., parts of hydrogen. The valency of the element is determined by the number of atoms of hydrogen to which the atom is equivalent.

3. With most solid elements it is the number which agrees with the law of atomic heats, that is, it is the number which, multiplied by the specific heat, gives an approximation to the constant 6·3.

4. It is the number which best agrees with the isomorphous relations of the element.

Name.	Symbol.	Atomic Weight.
Aluminium	Al.	27·
Antimony	Sb.	120·
Arsenic	As.	75·
Barium	Ba.	137·
Bismuth	Bi.	208·
Boron	B.	11·
Bromine	Br.	80·
Cadmium	Cd.	112·
Cæsium	Cæ.	133·
Calcium	Ca.	40·
Carbon	C.	12·
Cerium	Ce.	141·
Chlorine	Cl.	35·5
Chromium	Cr.	52·
Cobalt	Co.	59·
Columbium	Nb.	94·
Copper	Cu.	63·
Didymium	Di.	145·
Erbium	Er.	166·
Fluorine	F.	19·
Gallium	Ga.	69·

Name.	Symbol.	Atomic Weight.
Glucinum	Be.	9.
Gold	Au.	196.
Hydrogen	H.	1.
Indium	In.	113.5
Iodine	I.	127.
Iridium	Ir.	193.
Iron	Fe.	56.
Lanthanum	La.	139.
Lead	Pb.	207.
Lithium	Li.	7.
Magnesium	Mg.	24.
Manganese	Mn.	54.
Mercury	Hg.	200.
Molybdenum	Mo.	96.
Nickel	Ni.	58.
Nitrogen	N.	14.
Osmium	Os.	199.
Oxygen	O.	16.
Palladium	Pd.	106.
Phosphorus	P.	31.
Platinum	Pt.	195.
Potassium	K.	39.
Rhodium	Rh.	104.
Rubidium	Rb.	85.
Ruthenium	Ru.	104.
Scandium	Sc.	44.
Selenium	Se.	79.
Silicon	Si.	28.
Silver	Ag.	108.
Sodium	Na.	23.
Strontium	Sr.	87.5
Sulphur	S.	32.
Tantalum	Ta.	182.
Tellurium	Te.	128.
Thallium	Th.	204.
Thorium	Th.	234.
Tin	Sn.	118.
Titanium	Ti.	50.
Tungsten	W.	184.
Uranium	U.	239.
Vanadium	V.	51.
Ytterbium	Yb.	173.
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EXPERIMENTAL CHEMISTRY.

PART I. GENERAL PRINCIPLES.

CHAPTER I.

MATTER AND FORCE.

PUTTING on one side as unsolved, and probably for ever insoluble, the great metaphysical question whether the various objects and the complex phenomena we observe in nature have any real existence, or are merely images created in our own consciousness; and the farther and scarcely less recondite question whether, even if we admit the external existence of such objects and phenomena, we may not account for them without assuming the existence of any material substance; let us take the ordinary, and what may be called the common-sense view of the universe around us. In doing so, however, we must not venture to assert it dogmatically, or to stigmatise as absurd either of the other views above referred to, startling though they may appear. Both of them have been held by profound thinkers; neither can be proved to be erroneous, neither can be said to be in direct discord with known facts. The thoughtful student will find it good to revert to them from time to time as he advances in knowledge, and as he gains a clearer conception of their meaning he will find them less incredible than they appear at first.

Taking then the current view, we hold that all substances known to us, whether in the state of solid, liquid,

or gas, consists of *matter* which has an absolute existence. Matter is subject to the operation of certain agents called *forces*, which determine its position in space and its movements and properties. Whatever can produce or change motion is called a force. The forces can only be known to us by their operations on matter, just as motion cannot be conceived of except as of something moved. On the other hand, it is no less true that, apart from force, we can form no conception of matter.

Energy.—The power of doing work is called Energy. A body charged with electricity, a body hotter than its surroundings, and a moving mass are said to possess energy. Sometimes the energy is in action, as when a bullet is flying through the air, or an electric spark passing. In that case the energy is said to be *kinetic*. Sometimes the energy is quiescent, or passive, as when a cross-bow is bent, or an air-gun charged, then the energy is said to be *potential*. Potential energy is stored energy, and may at any time become active, or kinetic. The pulling of the trigger converts the potential energy of the cross-bow into kinetic energy. Gunpowder is a good example of potential energy.

Different Forms of Force.—The following are the chief kinds of force.

1. *Gravity.*—If a stone is suspended by a string, every one knows that the string is pulled by it. The stone and the earth attract one another, and if the string be cut, the stone will fly to the earth and come to rest. This kind of attraction is called gravitation, and it is said to be due to the force of gravity. It is found to exist in all forms of matter. Solids, liquids, and gases all tend to fly to the earth, and therefore all exhibit the force of gravity, though in very various degrees. Near its surface the great bulk of the earth causes its attraction to be so very much greater than that of any other body that the term gravity is generally applied only to the tendency of bodies to fly to the earth. In other words, gravity is generally taken as identical with *weight*. But it is right to remember that *all* bodies attract one another. One stone attracts another, one drop of water another drop; but the masses of the stones and the drops are so small

that the attraction is too slight to be noticed. That the force of gravity keeps the planets in their orbits, and regulates the movements of the heavenly bodies, is known to all. It is the most important doctrine of astronomy.

When the force of gravity is not resisted by some other force equal in power to itself, *motion* is produced.

The special case of gravity that we call *weight* is so important in the study of chemistry that it will be considered separately in the next chapter.

2. *Cohesion*.—The force which binds together the minute particles of which matter consists is called cohesion, but it is probably identical in nature with gravity. It varies very much in different substances, and even in the same substance in different states. To divide ice into smaller portions requires greater force than to divide water, and no force at all is required to divide steam. In solid bodies cohesion is stronger than in liquids; in gases scarcely a trace of it can be perceived. The *hardness* of a solid is the measure of the force of cohesion among its particles.

When cohesion is exerted between dissimilar substances it is often called *adhesion*. The use of glue is a good example of adhesion.

Experiments.—To prove that the force of cohesion exists in liquids, allow one pan of a pair of scales to rest on the surface of water. A number of weights must now be introduced into the other pan, before the first one will be removed from the water. When at last it rises, some water will still adhere to its under surface. The *adhesion* between the scale-pan and the water is not overcome, but only the *cohesion* between one portion of the water and the rest.

If only half the weights required in the above experiment be introduced, and a little ether be then poured on the surface of the water, the scale-pan will rise at once. The *adhesion* between the ether and the scale-pan is very slight.

Cohesion may also be illustrated by cutting a bullet in half, and pressing together with the hands the two clean and recently cut surfaces. If the cut has been well made, they will unite again with great force. Two pieces of

plate-glass will do the same, and the glass-makers have to be very careful not to allow large sheets of plate-glass to lie on one another.

3. *Heat* is an extremely powerful and important form of force. Its influence in chemistry is very extensive, and, like gravity, it will be considered in some detail in the next chapter.

4. *Light* is a force closely allied to heat, and is also closely concerned in many of the changes of chemistry. Almost the whole of the light and heat of our globe comes to us directly or indirectly from the sun.

5. *Electricity*.—*Experiment 1*.—Take a stick of sealing-wax, rub it briskly with a piece of flannel and hold it near some scraps of paper. They will be attracted by it and will fly upward and adhere to its surface. A tube of glass rubbed with silk will produce the same effect, but the substance called ebonite is more powerful than either. An ebonite comb that has been passed through dry hair will not only attract scraps of paper and small feathers, but in dry weather will emit a crackling sound and a torrent of small sparks, quite visible in a dark room.

In all these cases the strange and interesting force called *Electricity* is set in motion. The “electrical machine” is only a plate or cylinder of glass or ebonite, which, by turning a handle, can be made to rub against a silk cushion covered with amalgam. Electricity produced in this way is called *frictional electricity*.

Another kind of electricity is known, which is called Voltaic, or Galvanic Electricity. The following experiment will illustrate the way in which it is produced.

Experiment 2.—Take a tumbler two-thirds full of water, pour into it gradually an ounce and a half by measure of sulphuric acid, and fill the tumbler up with water. The mixture will become very hot, and must be allowed to stand until cold. Take a strip of sheet copper, and another of zinc, not too large to go to the bottom of the tumbler, and high enough to reach to the top. Brighten one of the small ends of each with sand-paper, and solder on each bright patch of copper wire about a foot long. It will improve the experiment very much if you can *amalgamate* the zinc plate; that is, coat it with mercury

(quicksilver). To do this the zinc must be cleaned by immersing it for a minute in the weak sulphuric acid, and then, while covered with the acid, pouring a little mercury on it and rubbing it over the surface with a small rag. The surface of the metal will then become as bright as a looking-glass.

The two metals are now to be immersed in the acid with the wires out. They should be about half an inch apart, and must not touch one another. They can be kept in their places by wedges of cork or wood. If the zinc plate has been amalgamated no action will take place *until the ends of the wires are made to touch one another*, when the zinc will immediately begin to dissolve in the acid and electricity to pass along the wires. Such an arrangement constitutes a simple and not very powerful form of *voltaic* or *galvanic* battery. It is called a "simple cell," and the ends of the wires farthest from the plates are called the "poles." If ten such cells are made, and are joined together by thick copper wires, the zinc of one cell to the copper of the next, long wires proceeding from the last zinc at one end and the last copper at the other, a tolerably powerful battery will be obtained, capable of decomposing water and doing other chemical work, but the arrangement is never a very powerful one, and it is better to buy or make one of the more powerful forms of the instrument.

If the two wires from the simple cell are brought to within half an inch of one another, and are then joined with a piece of fine platinum wire, the platinum will immediately become red-hot. With a more powerful battery a much greater length of wire may be heated to whiteness, and even melted by the electricity. Fine iron wire may be burned in the same way, but not quite so easily.

Magnetism is only a peculiar operation of electricity.

6. *Chemical Force, or Chemical Affinity.*—The last force that we have to consider is called chemical force, because the science of chemistry is almost entirely occupied with its nature and effects. Its operations are spoken of as chemical action.

Every one knows that iron, heated to redness, changes into black, brittle scales, and that exposed to moist air

or earth, it is converted into rust; that the expressed juice of the grape gradually turns to wine, and this again to vinegar; that wood in a stove, or oil in a lamp, disappears in burning; and that animal and vegetable substances in time putrefy, disintegrate, and finally disappear.

Iron scales and rust are iron altered in constitution: iron is hard, tenacious, of a greyish-white colour and brilliant; by heating to redness it becomes black, dull and brittle; on exposure to moisture it is converted into a powder of a yellowish-brown colour. Wine is altered *must*, in which nothing of the sweet taste peculiar to the grape-juice can be perceived; but it has acquired a spirituous flavour, together with a heating and intoxicating power, which was not in the *must*. Vinegar is altered wine; it has an acid smell and taste, and has lost its spirituous flavour, as well as its exhilarating properties, its tendency being rather cooling and sedative. Search much be made in the air for the oil and wood which have disappeared during combustion; both these substances are converted into vapour or gas, and heat and light are thereupon evolved with the phenomenon of fire. Of a similar nature are the changes which animal and vegetable substances undergo if kept for a sufficient length of time; they are gradually converted, as they putrify or decay, into various kinds of gas, some of which emit a very disagreeable odour.

Such processes, by which the weight, form, solidity, colour, taste, smell, and action of the substances become changed, so that new bodies with quite different properties are formed from the old, are called *chemical processes*, or *chemical action*.

Wherever we look upon our earth, chemical action is seen taking place, on the land, in the air, or in the depths of the sea. The hard basalt, the glass-like lava, become gradually soft, their dark colour becomes lighter, they crumble to smaller and smaller pieces, and are finally changed to earth. A potato placed in the earth grows soft, loses its mealy taste, becomes sweet and finally decays. The bud that sends forth a sickly pale shoot in a dark cellar, when exposed to the light and air grows up a vigorous, firm, and green plant, which, imbibing its nourishment from the moist air and soil, forms from their

elements new bodies, not to be found previously in the water or the air. A delicate network of cells and tubes pervades the whole plant, imparting to it firmness; these we call vegetable tissue, or woody fibre. We find in the sap, which passes up and down through these cells, albumin and other viscous substances; in the leaves and in the stalks, a green colouring matter—chlorophyll; and in the ripe tubers, a mealy substance—starch.

The potato forms one of our most important articles of food. The starch contained in it is not soluble in water, but when received into the stomach, quickly undergoes such a change that it can be dissolved or digested, and then introduced as a liquid into the blood. The blood comes in contact in the lungs with the inhaled air; the blood changes its colour, the air changes its constitution, and the heat which we feel in our bodies is developed. We must conclude, from these changes, that chemical action is going on in our own bodies.

If a piece of iron be heated to redness, till a thick crust of scales is formed around it, and then weighed, it will be found to have increased in weight; consequently, it must have been supplied with something ponderable from the air. This ponderable substance is a species of gas, called oxygen; but its union with the iron has become fixed, yet by other chemical processes it can be reconverted into its gaseous form. Accordingly, the crust consists of iron and oxygen, which have become most closely united with each other;—they are chemically combined.

The force which produces such changes as these is called **CHEMICAL FORCE**. In some books it is described as “chemical affinity,” because it is assumed that the tendency of substances to combine with one another is due to a kind of liking that they have for each other. Iron is so very fond of oxygen that it cannot help combining with it whenever it gets the chance. But it is unsafe in scientific matters to use names which assert more than we really know. Now we know that iron combines with oxygen, but we do not know that there is any “affinity” between the two. We know that force is exerted, but we do not know why. The conditions under which the force is exerted will be discussed in a future chapter.

Relation of the Forces to one another.—The forces we have been considering, unlike as they are in their manifestations, are very closely connected together. Almost any one of them may, under suitable circumstances, be made to produce another, and it has during the present century been shown that the one is directly evolved from the other. It has long been perceived to be impossible, under the present arrangement of the universe, for matter to be created or destroyed. It may change its form and pass ever so many times from one state of combination to another; but the total quantity remains always, as far as our means of observation enable us to judge, the same. The grandest generalisation of modern science is that which has taught us that energy is equally indestructible; that it also, like matter, may undergo various and countless changes; may appear, now as heat, now as electricity, now as chemical action, and now as visible motion, and yet its sum will always remain fixed and unalterable. There is no new creation of energy within the limits of our knowledge. It may lie hid for a time—for any time—but whenever we meet with any of its operations in nature we may confidently look for the cause in the disappearance of some previously active form of energy. The birth of one form of energy is always the death of some other, and the two are equivalent to one another in quantity.

This grand doctrine, which is known as the doctrine of the conservation of energy, requires complex apparatus, and great knowledge for its verification. The demonstration of it is indeed still incomplete, though it has gone far enough to satisfy every rational mind; but a few simple experiments will serve to illustrate the manner in which one force may give rise to others. In other words, we may prove the *correlation* of various forces, though the *conservation* of energy must be taken on trust.

Experiment 1.—Rub a piece of copper wire briskly with sand-paper for a minute or two. It will become so hot that the hand cannot bear to touch it, and it will readily ignite phosphorus or the tip of a lucifer-match.

Similar illustrations will occur to all. When the axle of a railway carriage gets dry, so much heat is produced

that the carriage is sometimes set on fire. A clever blacksmith can hammer a nail till it is red-hot, and so emits light as well as heat. During the boring of cannon the shavings are too hot to be touched, and, lastly, some savages are able to procure a light by rubbing two dry sticks together.

In all these cases there is a direct conversion of motion caused by force into heat, and even, in some cases, into light. Friction is but arrested motion—motion hindered or reduced by an opposing force. If the sand-paper had not pressed on the wire, the same amount of energy employed by the arms would have given rise to a much greater motion than that actually produced. The difference, though lost as motion, takes the new form of heat. When a pound weight falls to the earth from a height of 772 feet, its motion is lost, but the weight and the earth receive heat enough to raise the temperature of one pound of water one degree Fahrenheit.

We have already seen that friction may, under certain circumstances, produce electricity, so that we have gained illustrations of the conversion of motion into heat, light, and electricity. Let us now start from heat. The steam-engine affords the best possible illustration of the conversion of heat into *motion*. The work done by the engine is done by the heat expended, and so ultimately by the coals burnt. When heat becomes sufficiently intense, it is always accompanied by *light*, and Tyndall has devised a direct experiment in which invisible heat becomes visible light. In fact, heat and light only differ from one another as low and high notes in music do. The conversion of heat into *electricity* requires for its demonstration a delicate *galvanometer*, a somewhat expensive piece of apparatus. Heat is seldom directly converted into *chemical energy*, though the converse is one of the commonest of phenomena; but the development of the force of *cohesion* during the disappearance of heat can be readily demonstrated, as the following experiment will show.

Experiment 2.—Take a tumbler full of powdered crystals of sodium sulphate, and drench it with common hydrochloric acid. The salt will rapidly dissolve and

become liquid, and the force of cohesion will come into operation. But at the same time heat will be taken so rapidly from the materials in the tumbler that the outside will be covered with hoar frost, and if a little water in a test tube be plunged into the mixture it will freeze in a minute or two. The so-called freezing mixtures depend upon this principle.

Experiment 3.—Now try the reverse of this experiment. Take a clean Florence oil flask, nearly fill it with hot water, dissolve as much sodium sulphate in it as the water will dissolve (filtering the solution while hot if it appears dirty) and then, while the solution is nearly boiling, cork the flask tightly and allow it to get quite cold. There is now a great deal more of the solid in the solution than cold water could dissolve, but for some curious reason the excess does not immediately separate out. But take out the cork, and in a minute beautiful feathery crystals appear at the upper surface and pass downwards until the whole mass is solid. If the crystallization does not begin at once, drop in a small crystal of the sulphate, which will immediately produce the desired effect. As soon as the liquid has changed to solid, feel the outside of the flask with your hand. It is sensibly warm, showing that during the exertion and disappearance of the cohesive force heat has been given out. The experiment is even more easily made with the salt commonly called hyposulphite of soda.

The conversion of *electricity* into other forms of energy has already been illustrated to some extent in the battery. When the platinum becomes red-hot, it emits *heat* and *light*. When the poles of a powerful battery are connected with pieces of coke shaped like lead pencils, and the coke points after touching one another are separated by a short distance, a most intense light, called the “electric arc,” passes between them. The rays given out by it are similar to those of the sun, and, like the sun’s rays, they consist of three kinds:

1. Invisible heat rays.
2. Light rays.
3. Invisible actinic or chemical rays.

The invisible chemical rays are only to be recognised

by their effect in producing chemical changes. The interesting art of photography entirely depends on the power of rays to produce alterations of composition in certain chemical substances, and particularly the compounds of silver. The following experiment will illustrate this curious property.

Experiment 4.—In a room from which daylight is entirely excluded, but which may be lighted by a candle, dissolve sixty grains of silver nitrate (lunar caustic) in an ounce of cold distilled water. Pin a sheet of smooth white paper on to a flat board, and, by means of a flat camel's hair brush, paint it all over uniformly with the colourless silver solution. Allow it to dry in the dark; repeat the painting, and again allow it to dry. Paper so prepared is called "sensitive" paper. To show its properties, place upon it a piece of lace, a few fine feathers, or a sprig of fern, cover it with a sheet of window glass, and press the glass down with weights at the corners. The board may now be brought out of the dark room and exposed to sunlight, or to the strong diffused light of a bright day. The paper will soon begin to change colour, and after a time it will assume a rich brown tint. The whole arrangement must then be taken back to the dark room, the glass, ferns, &c., removed, and the paper immersed in clean rain or distilled water and well soaked, the water being changed a good many times. A beautiful image, in white on a dark ground, of the object employed will remain on the paper, which may now be dried and exposed freely to light, for the water will have washed away all the silver salt that has not been affected, and the picture will be permanent.

The waves of light consist of waves analogous to those of water, and the three kinds of waves just referred to differ from one another only in length, the length being measured from the crest of one to the crest of the next. The average wave-length for white light is about 1-50000 of an inch, and the eye perceives those only which, in round numbers, are between the 1-40000 and 1-60000 of an inch in length. Longer waves possess the energy of heat, and shorter waves of chemical action, but they do not affect the eye.

It has already been mentioned that the action of the sun's rays can also be produced by the rays from the electric light, and we have therefore an instance, to which many others might be added, of the production of chemical energy from electricity. Even with the single cell described in Experiment 2, page 4, some striking illustrations may be obtained, as will be seen from the experiments which follow.

Experiment 5.—Attach a small strip of platinum foil to the end of each wire. Cut a card so that it will stand upright in a wine-glass, and divide it in two. Then fill the glass with the card in it with a solution of the salt called potassium iodide, and bring one of the platinum *poles* into each of the two divisions which the card makes. The pole which is connected with the copper plate immediately produces a brown colour on that side of the card. The electricity breaks up, or *decomposes* the salt into two substances: potassium, which owing to its action on water is not seen, and iodine, the substance which gives the brown colour in this experiment. Pour into the brown liquid on one side of the card a little of the smooth, pasty liquid obtained by adding a good deal of boiling water to starch, stirring and cooling. The starch will combine with the iodine and form a beautiful blue colour.

Experiment 6.—Dissolve in some cold water as much copper sulphate (blue vitriol) as the water will take up, and substitute this solution for the potassium iodide used in the last experiment. The card is unnecessary. The salt will be decomposed, and copper will be deposited on the platinum pole which proceeds from the zinc plate. If the poles are reversed the copper will soon dissolve off the one platinum plate and appear on the other. This experiment is an example of the process called electroplating. Other metals, such as gold and silver, may by analogous means be deposited from solutions. The decomposition of liquids by electricity is called *electrolysis*.

Experiment 7.—*Decomposition of Water by Electricity.*—With the assistance of a more powerful battery, and the little piece of apparatus shown in Fig. 1, water may be decomposed into the two gases of which it consists. Thick

platinum wires pass through the sides of a glass and terminate inside in flat plates (stout platinum wires hammered flat at the ends do very well). The holes in the glass can easily be made by a bradawl filed to a square point and constantly moistened with a mixture of turpentine and camphor, and the wires may be cemented in the holes with sealing-wax. The glass is filled with water to which a drop or two of sulphuric acid has been added, and over each of the plates a test tube filled with the same acidulated water is inverted. The two ends of the wires are now connected with the poles of the battery. In an instant bubbles of gas rise from the two plates into the test tubes. The gas which comes from the pole connected with the zinc is called *hydrogen*, that from the other *oxygen*; and it will be observed that the volume of the former is twice as great as of the latter. When a little gas has been collected, the mouths of the tubes may be closed with the thumb, the tubes removed, and the gas examined with a taper in the manner described under their respective names.

Fig. 1.



This is the most important instance of the production of chemical energy at the expense of electricity that we possess. Electricity disappears during the experiment, but the hydrogen and oxygen which are produced are able, as we shall find hereafter, to exert an enormous force when they once more combine together.

Lastly, let us study the forces which are called into existence during the exertion and consequent disappearance of *chemical* energy. We have, in point of fact, just been considering one of the most interesting cases of this kind in the voltaic battery. In the battery, the electricity which is called into being proceeds directly from, and is proportionate in amount to, the chemical energy which is destroyed in the cell. The zinc is constantly being dissolved by the sulphuric acid, and the force which is thereby exerted immediately takes the new form of electricity. From this electricity we can, as we have seen, produce heat, light, and finally, chemical force, so that we can ultimately recover the very same form of force that we lost in the battery. One of Faraday's

greatest discoveries proved that there is no loss of force throughout these changes, but that, provided we can prevent the force from assuming other forms, the chemical energy produced by the decomposition of the water is exactly equal to the energy lost in the battery.

One or two simple experiments will illustrate the direct production of other forces from the chemical, but it is unnecessary to multiply them, because the whole study of chemistry is full of such illustrations.

Experiment 8.—Take a piece of phosphorus, about the size of a small pea, and about an equal quantity of iodine. Place them side by side on a slate or a piece of tin-plate, and push the phosphorus with a knife till it touches the iodine. Chemical action is instantly exerted, and the mass bursts into flame, producing thereby *light* and *heat*.

Experiment 9.—Warm two tumblers: put into one a few drops of strong ammonia, and into the other a few drops of hydrochloric acid. Each will give out a gas, and when the mouths of the tumblers are held together these gases will combine, and a dense white smoke will be produced, which will before long settle down on the inside of the glasses in the form of a white, solid powder (sal-ammoniac). The chemical action is in this case accompanied by the development of the force of *cohesion*. The lost force takes the form of heat, which, however, can only be perceived by the thermometer.

When gunpowder or any other explosive substance is burnt, it is converted into gas; very energetic chemical action takes place, and a great deal of force is exerted. But the force is entirely reproduced, for the cohesion of the solid is destroyed, a great expansion, which is nothing but a motion of the particles, takes place, and heat and light are evolved. Every one knows what powerful work the expansion may be made to do. Heavy shot can be thrown from a gun to an enormous distance, and solid rocks torn in fragments by the burning even of a small weight of the powder.

The result of the researches of the present century has been to convince scientific men that, various as are its manifestations, there is in nature but one force, namely, **MOTION**: motion of masses, motion of particles, motion in

lines, motions of rotation and vibration, and motion in other modes. As regards heat and light this strange doctrine may be said to be already demonstrated, for these remarkable forces have been proved to be merely peculiar modes of motion in the particles of matter. And the close connection of heat with the other forms of force forbids us to doubt that its nature is essentially similar to theirs, and that they also are modes of motion. Science has still a great work before her in the investigation of force, but enough has already been accomplished to enable us to take a far wider view of the kingdom of nature than was possible to our fathers, and to reveal to us something of the unity and simplicity that underlies the marvellous complexity of the universe.

OBJECTS OF CHEMICAL INQUIRY.

It has already been mentioned that the science of chemistry is concerned with the operations of chemical force. There are *four chief points* to which the attention of the chemist is mainly directed in his study of the solid, liquid, and gaseous substances which are met with in nature.

1. *Their composition.*—Take a piece of bone. How is it affected when strongly heated in a furnace? It becomes whiter, lighter, and less solid than before (bone-ash). But how is it affected when heated in a covered vessel? It becomes lighter, and black (bone-black). If exposed to boiling water, or to steam, how is it affected? It becomes lighter, and remains white; but in the water is dissolved gelatine. How in hydrochloric acid? It becomes transparent; the bone earth is dissolved, and a gristly mass remains, which, when boiled with water, turns to gelatine. What is the action of fire upon the gelatine? In a covered vessel it is converted into charcoal, in an open one it burns and disappears. These few experiments show that the bone contains gelatine which is combustible, and an earth which is not so; they show at the same time that it is the carbonized gelatine which, in the second experiment, colours the bone-earth black, and makes it bone-black; that this gelatine dissolves in water, but not in hydrochloric acid, &c. Gelatine and bone-earth are called the *proximate constituents*

of bone, but by continued chemical processes these can be resolved still further, that is, separated into simpler constituents. In bone-earth are found phosphorus, a metal (calcium), and oxygen; in the gelatine, besides carbon, three other bodies—oxygen, hydrogen, and nitrogen. These bodies can be decomposed no further by any known method of analysis, and are therefore called *simple bodies*, or *elements*. There are now about seventy known elements, and almost every year adds to their number. This separating of compound bodies into simple ones is designated by the name of *decomposition*, and the process of ascertaining the composition of any substance is called *analysis*.

When a substance contains two or more elements, held together by chemical force, it is called a *compound*. Compounds are always quite different in properties from their constituents.

When a substance contains two or more other substances (elements or compounds), not held together by chemical force, but present as it were accidentally together, it is called a *mixture*. In a mixture the properties of the separate ingredients are still perceptible.

This distinction between element, compound, and mixture is very important. The following illustrations will assist to fix it in the memory. Water is a *compound* of the *elements* hydrogen and oxygen, which are held together by chemical force. Air is a *mixture* of the elements oxygen and nitrogen, which are not in chemical combination with one another. Gunpowder is a *mixture*, containing the *elements* carbon and sulphur and the *compound* saltpetre. The *compound* saltpetre contains the *elements* potassium, nitrogen, and oxygen.

2. *The changes which they undergo by the action of other bodies and of the various forces upon them.*—Phosphorus, which is obtained from bones, is luminous in the air, and is gradually converted into an acid liquid; it unites with the oxygen of the air and with water, as iron does when exposed to moist air. If the phosphorus is gently heated, this union is attended with a vivid combustion, and there is formed a body which is different from the former; to which, if water and lime be added, a new body is formed,

very similar to bone-ash; it is, in fact, artificial bone-ash. The number of new bodies which may be produced by the union of the elements with each other, or with compound bodies, is infinite; and entirely different substances are often formed, according as the combination takes place under the influence of cold or heat, in water or in air, in greater or smaller quantities. This is *combination* or *synthesis*.

3. *The causes of chemical changes and the laws according to which they take place.*—If chemical experiments are performed as they should be, with the balance at hand, it will soon be observed, that when two different bodies which can unite with each other are brought together, sometimes a part of the one, sometimes a part of the other, remains free. Further experiments will show how much of one body, in weight, can be united with the other. If all bodies are tested in the same manner, the certainty is finally attained, that all chemical combinations take place only in fixed, unchangeable proportions, and that to every individual body is assigned a definite weight, with which it always enters into any combination whatever. (Chap. III.) This certainty is called a *natural law*. Many such laws of nature have already been ascertained, and they serve as a certain guide to the chemist in his labours, since they cannot, like human laws, be arbitrarily evaded or changed. By them alone we attain to a scientific insight into chemical processes, and to the capability of putting direct questions to bodies by experiment, and of testing the truth of the answers received. An explanation of chemical processes based on natural laws, which presents a clear idea of the subject to the mind, is called a *Theory*.

4. *The extent to which the facts which have been discovered may be made useful to man.*—When the chemist discovers a new body, or a new property in one already known, or a new method of synthesis or analysis, he imparts his discovery to the apothecary, the physician, the farmer, the manufacturer, and the tradesman, that experiments may be instituted for the purpose of ascertaining whether any advantage, facility, or improvement can be derived for pharmacy, medicine, agriculture, or the arts. Phos-

phorus ignites spontaneously at a gentle heat; it is used in lucifer matches. Taken into the stomach, it acts as a violent poison; it is at present one of the most common means for the destruction of vermin. The constituents of bone-earth and those of gelatine have been found to be universally present in the seeds of different kinds of corn; the chemist concludes from this that pulverised bones should yield an excellent manure; the agriculturist demonstrates this by experiments on a large scale. In bone-black the property has been discovered of attracting many substances held in solution in liquids, and of condensing them in itself; on account of this property it is used for making impure water potable; the sugar-refiner employs it to make brown syrup colourless. This is *applied or practical chemistry*.

Nothing is better calculated to excite an interest in chemical knowledge than a consideration of the useful application which can be made of it in every-day life. Chemistry teaches the apothecary how to compound and prepare his medicines; it teaches the physician how to cure maladies by means of these medicines; it not only shows the miner the metals concealed in rocks, but aids him also in smelting and working them. Chemistry, in connection with physics, has been the principal lever by which so many arts and trades have been brought to such a degree of perfection within the last few decades, and by its means we have been supplied with numberless conveniences of life that were not enjoyed by our fathers. It is obvious that the farmer must regard chemistry as his indispensable friend, for it is this alone which acquaints him with the constituent parts of his soil, with the proper nutriment of the plants he wishes to cultivate, and with the means whereby he can enhance the fruitfulness of his fields.

The ancient so-called Elements.—The dogma of the ancient philosophers, that there are four elements, earth, air, water, and fire, has long been known to be erroneous, although the words are still repeated very often. We now know that three of them are either mixtures or compounds of the true elements, and that fire is but a peculiar operation of heat and light.

CHAPTER II.

GRAVITY, HEAT, PRESSURE.

GRAVITY,

THE force of gravity is chiefly interesting to chemists because it confers upon all kinds of matter the property called weight.

WEIGHING AND MEASURING.

Weighing.—The balance is the plumb-line, as well as the touchstone in chemical experiments; it teaches us how to ascertain the true composition of bodies, and shows us whether the questions put, the answers received, or the conclusions drawn from them, are correct or false. Hence it cannot be too strongly recommended to those commencing the study of chemistry to use the balance even in simple experiments. For the experiments described in this book, a common apothecaries' balance is all that is requisite.

Avoirdupois weight is most frequently used in England. The only numbers required in chemistry are the following:—

Pound.		Ounces.		Grains.
1	=	16	=	7000
		1	=	437.5

Apothecaries' weight is somewhat different, but the grain is the same.

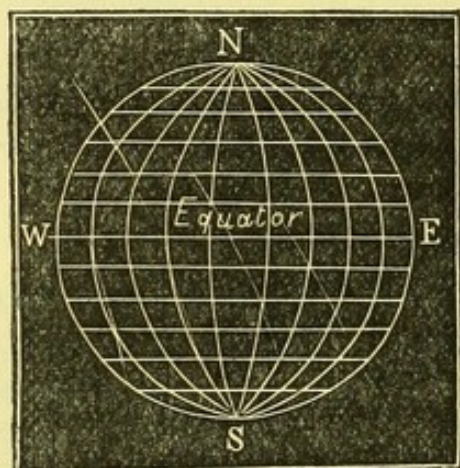
Pound.		Ounces.		Drachms.		Scruples.		Grains.
1	=	12	=	96	=	288	=	5760
		1	=	8	=	24	=	480
				1	=	3	=	60
						1	=	20

The French system of weights and measures, which is now almost universally adopted by chemists, is charac-

terised by great simplicity, all its divisions being made by ten; hence the name *decimal* weight and measure. Its unit is derived from the size of our globe.

In order to define the different localities on this globe, imaginary circles, as is well known, have been drawn around it. Those which pass round the earth from east to west, the largest of which is the equator, are called parallels of latitude; those which pass round the earth

Fig. 2.



lengthwise, intersecting at the poles, *meridians*. The parallels of latitude gradually become smaller towards the poles; the meridians, on the contrary, are all of equal size. The circle N E S W represents a meridian or circle of longitude. The fourth part of this circle, or, what is the same thing, the fourth part of the circumference of our earth, as N E, is the basis of the French system.

This quadrant was divided into ten million parts, one of which was taken as the unit, under the name of *metre*. A metre is nearly forty, or about $39\frac{3}{8}$ inches in length. The smaller measures are produced by dividing by ten, and are designated by Latin prefixes; the larger ones by multiplying by ten, and are designated by Greek prefixes.

Smaller Measures.		Larger Measures.	
Metre.		Metre.	
Decimetre	= $\cdot 1$ or $\frac{1}{10}$ metre.	Decametre	= 10 metres.
Centimetre	= $\cdot 01$ or $\frac{1}{100}$ "	Hectometre	= 100 "
Millimetre	= $\cdot 001$ or $\frac{1}{1000}$ "	Kilometre	= 1,000 "
		Myriametre	= 10,000 "

The millimetre is nearly $\frac{1}{25}$ of an inch.

The system of weights are derived from the measure of length, in the following manner. A cube of water at its greatest density (at the temperature 39° F.) (4° C.) was taken as unit, which would measure one centimetre in each direction, and the weight of this quantity of water was called a *gramme*. This is taken as the unit of the decimal weights, and is multiplied or divided by ten.

Smaller Weights.		Larger Weights.	
Gramme	= 1.0	Gramme	= 1
Decigramme	= .1 or $\frac{1}{10}$ gramme.	Decagramme	= 10 grammes.
Centigramme	= .01 or $\frac{1}{100}$ "	Hectogramme	= 100 "
Milligramme	= .001 or $\frac{1}{1000}$ "	Kilogramme	= 1,000 "
		Myriagramme	= 10,000 "
One gramme is equal to 15.4321 grains.			
One kilogramme is equal to 2.204 lbs. Av.			

It is well enough known that the body whose weight is to be ascertained must be put into one scale, and in the other, weights sufficient to restore the index to its original perpendicular position. The weight of a body thus determined is called its *absolute weight*. Thus, a piece of sugar weighing two ounces has an absolute weight of two ounces.

Measuring of Liquids and Gases.—The English system for liquids is as follows:—

Gallon.	Quarts.	Pints.	Fluid Ounces.	Fluid Drachms.	Minims.
c.		o.	fl. $\frac{3}{4}$.	fl. $\frac{3}{4}$.	℥.
1	= 4	= 8	= 160	= 1,280	= 76,800
	1	= 2	= 40	= 320	= 19,200
		1	= 20	= 160	= 9,600
			1	= 8	= 480
				1	= 60

1 gallon of water at 60° Fah. = 277.276 cubic inches, weighs 10 lbs. Av. = 70,000 grains.

1 fluid ounce of water weighs 1 ounce Av. = 437.5 grains.

Grains of water are likewise frequently employed as a measure for liquids. 1,000 grain measures = $\frac{1}{70}$ of a gallon.

Gases are generally measured in cubic inches.

The standard of measure in the French system is the *litre*, which is the same as one cubic decimetre, or a cube of nearly 4 inches on each side. The same prefixes are used for the fractions and multiples as for the metre and gramme, but it will be seen that the litre is also equal to 1000 *cubic centimetres*. Cubic centimetres are generally used in chemistry for the measurement both of liquids and gases. As one cubic centimetre of water weighs one gramme, a litre of water weighs 1 kilogramme. It is almost exactly equal to $1\frac{3}{4}$ pint.

The student will find it advantageous to use the French weights and measures. They are now generally employed by scientific men.

SPECIFIC GRAVITY.

Ice floats in water, iron sinks in it, because the former is lighter, the latter heavier, than water. But if we put a piece of ice in spirit it sinks, or if we put a piece of iron upon mercury (quicksilver), it floats; consequently, ice is heavier than spirit, iron lighter than mercury. It also follows that spirit is lighter than water, since it can support less weight, and mercury heavier than water, as it can bear a greater weight. The terms heavier and lighter, in this sense, correspond to what in scientific language is called *specifically heavier* or *specifically lighter*, and *equal bulks* are always to be understood in speaking of the comparative weights of bodies. The expression, *ice is lighter than iron*, means, therefore, that, taking equal bulks of each, the former weighs less than the latter; and when we say that mercury is heavier than water, we mean that in equal volumes, as a pint, for instance, the mercury has a greater weight than the water. But in absolute weight, no regard is paid to the volume of substances.

In order to ascertain *how many times* heavier mercury is than water, or iron than ice, it is only necessary to weigh equal volumes or portions of each, and to compare their weights at the same temperature. If, for example, we take five vessels, each of which would contain exactly 100 grains of water, and fill them respectively with spirit, ice, water, iron, and mercury, the following differences in weight will be found: the vessel with spirit would weigh 80 grains; with ice, 90 grains; with water, 100 grains; with iron (if quite pure), 780 grains; with mercury, 1360 grains.

To facilitate the comparison of the numbers which denote how much greater the specific gravity of one body is than that of another, water has been fixed upon as the standard or unit. Therefore, in the above case, the question is, how much lighter than water are spirit and ice, and how much heavier are iron and mercury? or, in other words, how many times is 100 contained in 80, 90, 780, and 1360? The other numbers, then, are to be divided by 100, the weight of water, and there is found for

Spirit, $\frac{80}{100}$, or, in decimals, 0.80; it is therefore $\frac{1}{5}$ lighter than water.

Ice, $\frac{90}{100}$, or, in decimals, 0.90; it is therefore $\frac{1}{10}$ lighter than water.

Iron, $\frac{780}{100}$, or, in decimals, 7.80; it is therefore $7\frac{8}{10}$ times heavier than water.

Mercury, $\frac{1360}{100}$, or, in decimals, 13.60; it is therefore $13\frac{1}{2}$ times heavier than water.

These numbers represent the *specific weights* (sp. gr.). Thus, according to calculation, spirit having a specific gravity of 0.80, 80 parts of it would occupy the same space as 100 parts of water; therefore it is only four-fifths as heavy as water, or, what is the same thing, one-fifth lighter than water. The specific gravity of mercury being 13.6, that is, $13\frac{1}{2}$ parts of mercury do not take up any more space than one part of water; since it is about $13\frac{1}{2}$ times heavier than water.

Determination of Specific Gravity.—Experiment.—To determine the specific gravity (the density) of a fluid, a bottle is weighed, then filled with water, and again weighed. This gives the weight of the water. Now pour out the water, and refill the bottle either with spirit, syrup, lye, beer, or some other liquid, and ascertain by the balance the weight of each. Then divide the weight of each of these fluids by the weight of the water, and the quotient indicates the specific weight. It is very convenient to use a bottle made to contain exactly 100 grammes of water. As bodies expand by heat, and so become specifically lighter, they must be compared at a constant temperature.

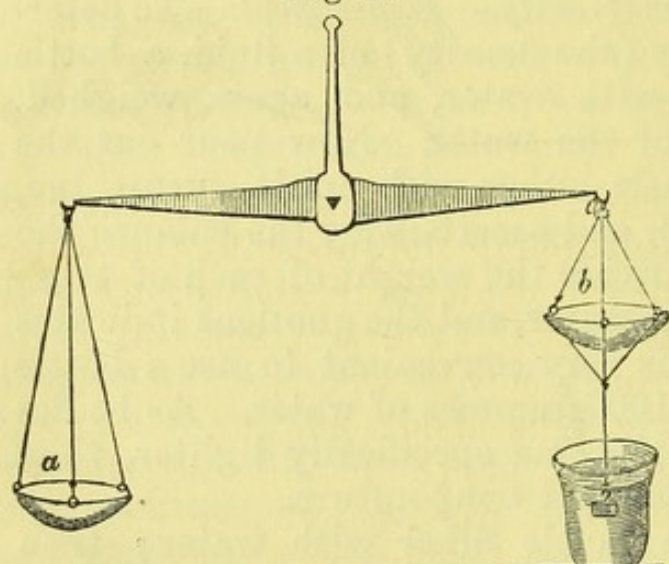
Experiment.—Weigh a bottle filled with water; then place a 20-gramme weight on the pan which holds the weights, and by the side of the bottle nails enough to adjust the beam. Remove both the nails and the bottle from the pan, and put the nails into the bottle. A bulk of water will be displaced equal to that of the nails; to determine its amount, replace the bottle, after it has been thoroughly wiped on the outside, upon the pan, and remove weights from the other pan until the equipoise is restored. The weights taken away (about 2.6 grammes) form the divisor, and the 20 grammes the dividend; the quotient, $\frac{20}{2.6} = 7.7$, is the specific gravity of iron of which the nails were made. English weights may, of course, be used.

The specific gravity of gases is determined by a process

which is similar in principle to the above, but is much more difficult. The standard of specific gravity for gases now generally used is hydrogen, which is the lightest gas known. When we say that the specific gravity of oxygen is 16, we mean that it is 16 times heavier than hydrogen. Air is sometimes taken as the standard. It is 14.5 times heavier than hydrogen, so that, knowing the specific gravity of a gas in relation to hydrogen, it is easy to find it in relation to air by dividing by 14.5. Thus the specific gravity of oxygen as compared with air is $\frac{16}{14.5} = 1.1$.

Experiment.—If we have to determine the specific gravity of a piece of iron, or of any other body which cannot be put into a bottle, it must be fastened by a piece of fine thread to the pan of a common balance (Fig. 3, *b*), the cords of

Fig. 3.



this pan having been previously shortened. Weigh the body first in air, and then in water, immersing it an inch deep. When it is immersed, the opposite pan falls; consequently iron must be lighter in the water than in the air. If the iron in the air weighed 240 grains then, in order to re-

store the equilibrium, it will be necessary, as in the former experiment, to remove from the pan *a* 32 grains, equal to the weight of the bulk of water displaced by the iron. The loss of weight is the same, whether the water be removed from the vessel or merely displaced within it. This forms the divisor, and 240, the weight of the iron in the air, the dividend, giving the quotient $\frac{240}{32} = 7.5$.

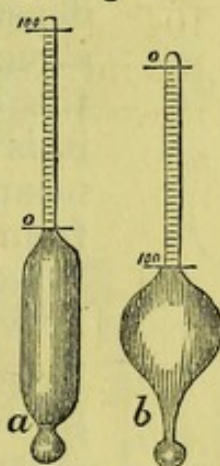
Every substance becomes as much lighter in water as the quantity of water displaced weighs. If it displaces less water than its weight in the air, it sinks; if more, it floats. Even very heavy bodies can be made to float by increasing their volume; ships are constructed of iron, although it is eight times heavier than water; a tumbler

floats upon water, and yet the specific gravity of glass is three or four times greater than that of water. A thick piece of iron, weighing half an ounce, loses in water nearly one-eighth of its weight; but if it is hammered out into a plate or vessel of such a size that it occupies eight times as much space as before, it then loses its whole weight in water, and will float, sinking just to the brim. If made twice as large, it will displace one ounce of water—consequently twice its own weight. It will then sink to the middle, and can be loaded with half an ounce weight before sinking entirely.

Hydrometer or Areometer.—The same floating body will sink to a greater or less depth in different liquids—deeper in the lighter ones, and not so deep in those which are denser. This has suggested a very convenient instrument for determining the specific gravity of liquids—the *hydrometer*, or *areometer*. This instrument consists of a hollow

glass tube, made as represented in Fig. 4. The interior is hollow, and blown out into a bulb at the lower end, to cause it to float; the under part is loaded with quicksilver or shot, to give it a vertical position. The narrow tube serves to denote the depth to which it sinks in any liquid, by means of a scale of degrees, with which it is furnished. There are various instruments of this kind, especially adapted for determining the density of spirits, brandy, oil, soda, syrup, &c. If a hydrometer for weighing spirits is put into water, it sinks only to the lowest point on the scale, 0° (Fig. 4, *a*); but in the strongest alcohol, which is much lighter than water, it sinks to the highest point, 100° . A scale for testing soda (Fig. 4, *b*) must, on the contrary, have the 0° point at the top of the scale, to which it would sink in pure water; for soda being heavier than water, the instrument would be more or less buoyed up in it, according to its strength. In hydrometers for lighter liquids, the degrees proceed from the bottom to the top; in those for heavier liquids, from the top downwards. In some of these scales the degrees are arbitrary, and in order to convert them into the corresponding specific numbers,

Fig. 4.



tables, constructed for the purpose, must be referred to.

Experiment.—Pour brandy into a cylindrical jar, and observe the degree which it marks on the hydrometer; then put it in a warm place, and when lukewarm, again note the degree, which will be higher than before, as the heat has expanded the liquid, made it lighter, and consequently apparently stronger than it really is. The specific gravity of all bodies, when warmed, is less than when cold. On this account, in determining the density of bodies, regard should be paid to their temperature, and it is usual to consider 60° F. (15.5° C.) as the mean temperature.

In some hydrometers the mercury serving as the counterpoise has been ingeniously contrived also to indicate the degree of heat of the liquid, by connecting with it a graduated tube. The small scale, *a* (Fig. 5), denotes the temperature, the long scale, *b*, the density. The small scale is frequently so constructed that the degrees correspond to those on the long scale, and in order to guard against error it is only necessary to add the degrees below the mean temperature to the density, or to subtract from the density those above.



Gold is nineteen times, and silver ten times, heavier than water; gold alloyed with silver must, therefore, have a less specific weight than pure gold. The specific weight of brass is only = 8. Alcohol and ether are lighter in proportion to their purity and strength, while lye, syrup, the acids, &c., increase in density according to their purity. Hence it is evident how important it is, in many cases, to know the specific gravity of a body in order to judge of its quality.

HEAT.

We have already seen that heat is often given out during chemical action. When chemical force is exerted, heat is generally, if not always, produced; and, on the other hand, many chemical changes will not take place without the assistance of heat.

The artificial heat which is employed in chemical operations and in ordinary life is obtained from chemical action. Coal, wood, spirit, oils, and all the other kinds of fuel undergo a chemical change when they burn, which will be fully explained hereafter. They combine with the oxygen of the air, and in doing so chemical force is consumed and changed into heat.

Sources of heat for chemical operations.—If the student is fortunate enough to have gas at his command, he will seldom need any other fuel.

The gas may be brought from the nearest burner, from which the jet has been screwed off, by means of an india-rubber tube. For most purposes, the burner known as the Bunsen burner (Fig. 6) is convenient. The gas becomes mixed with air in this burner before burning, and burns therefore with a smokeless flame, which gives very little light, but an intense heat. Little iron caps, *a* (Fig. 6), are sold with each burner, which fit on it, and convert the single flame into a ring of smaller jets. Some of the burners made by Fletcher of Warrington are exceedingly useful in chemistry, and they are very cheap.

If gas cannot be obtained, a small spirit-lamp, such as shown in Fig. 7, may be used for ordinary purposes, and methylated spirit, which is very cheap, may be burned in it. More powerful spirit-lamps are, however, sold, and small oil-lamps may sometimes be employed for the sake of economy when the heat has to be employed for a long time. Furnaces are scarcely necessary to the beginner in chemistry. The few crucible operations he has to perform may be done in a common fire.

Fig. 6.

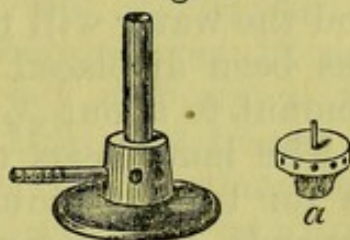
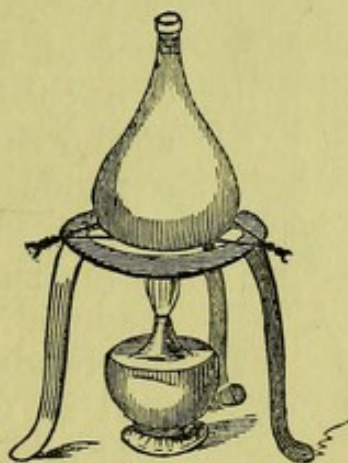


Fig. 7.



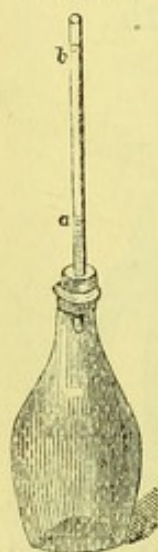
EFFECTS OF HEAT.

Expansion of Liquids.—*Experiment 1.*—Counterpoise a flask—that is, place it on one of the pans of a balance, and equipoise it by weights or shot put into the opposite

pan; then fill it to the brim with water, and note the weight of the latter. Warm the flask on a tripod over a spirit-lamp or gas-burner. The water will expand, and part of it will run over, and may be carried away by a gutter of tin plate just under the neck of the flask. When it begins to boil, remove the lamp, and let the vessel cool, and the water will then sink below the neck. How much has been displaced is found by its loss in weight; it will amount to about $\frac{1}{22}$ of the first weight.

The lamp heats the bottom of the glass vessel, which in its turn communicates heat to the water. The heat expands the water, consequently it occupies a greater space than before, and part of it must run over. Hence it follows that warm water must be lighter than cold water. If a pitcher filled with two pounds of ice-cold water be afterwards filled with boiling water, it will weigh about an ounce and a half less.

Fig. 8.



The same occurs with all other liquids, and, indeed, also with solids and gases; hence, it may be stated as a natural law, *that all bodies expand by heat, and contract on cooling*. But the amount of expansion is very different in different bodies at the same temperature; alcohol, for example, expands much more, mercury much less, than water. When fluids are to be bought and sold by measure, an advantageous application may be made of this principle. If a hundred measures of brandy or alcohol were purchased in hot, and sold in cold weather, there would be a loss of four or five measures; therefore we should gain by buying in winter and selling in summer.

Experiment 2.—In order to observe more accurately the expansion of water by heat, adapt to a flask a cork, rendered so soft by gently striking it with a piece of wood that it may be exactly fitted to the opening by mere pressure; perforate the cork with a round file or cork-borer, and make the hole just large enough to admit a glass tube tightly. Fill the flask with water, so that, when the cork is firmly pushed in, the water shall stand at about *a* (Fig. 8), and heat it as in the former experiment.

The water, which in the former experiment was displaced

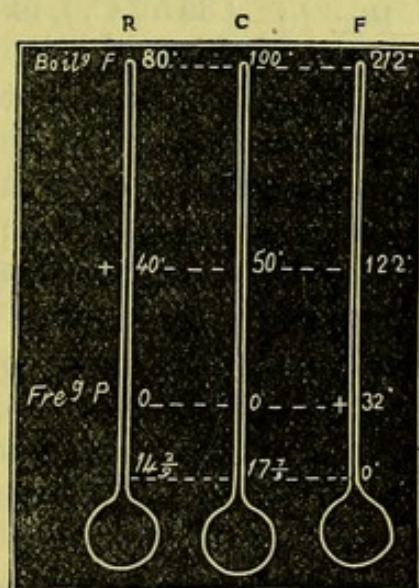
from the flask, now rises in the tube, and this higher in proportion to the smallness of its bore. By this means very slight changes of space are rendered visible, and these deviations may be applied to the *measurement* of heat. This is done by particular instruments called thermometers.

Thermometer.—Water might be employed for measuring heat, by marking the boiling and freezing points, and graduating the intervening space; but mercury is far better adapted to the purpose, as it boils and freezes at greater extremes of temperature, and becomes hot and cold more quickly; hence it more quickly denotes the variations of heat and cold: moreover it expands more uniformly than water.

The vessel containing the mercury may also be regarded as consisting of a flask and tube, but which, instead of being joined by a cork, are composed of one entire piece.

Having introduced into it a sufficient quantity of mercury, and sealed the open end by fusion, it is immersed in melting ice, and the point to which the quicksilver falls is marked *freezing point*; that to which it rises in boiling water, *boiling point*. The space between these two points can now be divided into degrees, to form the *scale*. The degrees below the freezing point must be of the same dimensions as those above. There are several scales in use, though it is to be regretted that more than one

Fig. 9.



has been adopted. The most common are the three following:—*Reaumur's* (R.), divided into eighty degrees; the Centigrade of *Celsius* (C.), into one hundred; and *Fahrenheit's* (F.), into one hundred and eighty degrees. The difference between these can be easily seen in the annexed figure.

According to R. water freezes at 0° and boils at 80°; according to C. it freezes at 0°, and boils at 100°; according to F. it freezes at +32°, and boils at 212°.

Fahrenheit, a philosophical-instrument maker, commenced reckoning very strangely, not at the freezing point, but at 32° below it. His scale is still in common use in England, and the high numbers found in English works are thus accounted for. In Germany, *Reaumur's* thermometer is used, except for scientific purposes, when the *Centigrade*, in common use in France also, is employed. In order to compare these thermometers with each other, it need only be remembered that 4° R. are equal to 5° C. or 9° F. In reducing the degrees of Fahrenheit's scale above 32° to those of the Centigrade, the number of degrees above or below 32° must be multiplied by 5 and the product divided by 9. And to reduce those of Fahrenheit to Reaumur, the number of degrees above or below 32° must be multiplied by 4 and divided by 9. To the degrees above 0° , the sign + is prefixed or understood; to those below, the sign - is invariably added.

A cylindrical thermometer, graduated to about 600° F. (315° C.), is best suited for chemical experiments, as shown in the annexed figure, because it can be easily adapted to a perforated cork, and then fitted to a flask, in which liquids are to be heated to a certain temperature. The degrees above the boiling point are to be divided off at distances equal to those below.



Mercury freezes at -38° F. (-39° C.). In the northern regions of the earth a degree of cold of -58° F. (-50° C.) has been observed, and by artificial means the temperature can be lowered to temperature lower than -148° F. (-100° C.).

When great degrees of cold are to be measured, alcohol is used in the construction of this instrument, as it does not congeal at -148° F. (-100° C.).

Mercury boils at 675° F. (357° C.); therefore its use must be limited to temperatures below this point. The high temperatures attending ignition are sometimes measured by the expansion of bars of platinum, a metal which does not melt in any ordinary furnace. Such an instrument is called a pyrometer. The determination of very high temperatures is, however, obtained by other means.

Expansion of Solids.—If an iron vessel, when cold, is just large enough to pass through the door of a furnace, it cannot be removed from it when heated. The iron bands or tires of carriage wheels are applied while hot to the frame, and on cooling they contract and bind the wood-work together with great force. A metallic disk, which, when red-hot, fits exactly into a circular box, will, on cooling, become loose, and shake in it. The tire and the disk both become smaller on cooling. These examples show that solids also are expanded by heat and contracted by cold, and explain many of the phenomena of common life. Clocks are apt to go faster in winter, and slower in summer, because the pendulums elongate in summer, and, consequently vibrate slower; while in winter they become shorter, and vibrate more rapidly. A piano gives a higher tone in a cold than in a warm room, on account of the contraction of the strings. A nail driven into the wall becomes loose after a time, because the iron expands in summer and contracts in winter more than the stone or the wood, and thus the opening is gradually enlarged. For this reason, in the construction of railroads, the rails must not be laid too close together; in the arrangement of steam-pipes, these must not be too firmly inclosed; in roofing, the zinc plates, instead of being nailed *together*, must overlap each other, that they may neither tear nor warp on alternate contraction and expansion.

Brittle bodies, as glass and porcelain, expand or contract so rapidly, by sudden heating or cooling, that they break.

Experiment 3.—Wind round a bottle two bands of paper, *a* and *b*, Fig. 11, and secure them firmly with thread; pass a piece of string round the bottle, between these folds of paper, and move the bottle quickly to and fro on the string until the latter breaks. Then immediately pour cold water upon the place, and the glass will break as evenly as if cut. The sharp edges can be removed with a file. In this manner, common white or green glass bottles may be converted into vessels adapted to chemical and other purposes.

It is well known that heat is produced by the friction of

Fig. 11.

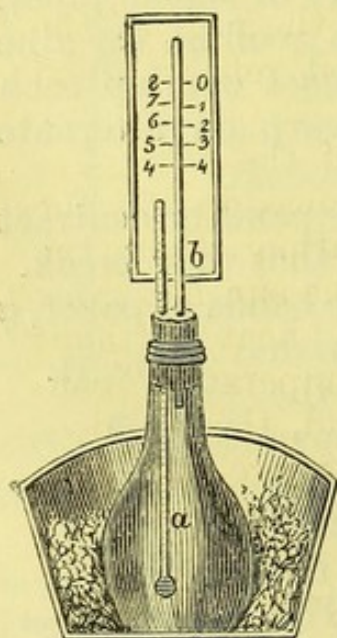


two bodies upon each other; that by sliding quickly down a line or a pole by the hands, these will be burnt; and that rapid motion will ignite the axles of a carriage, unless they are well greased. Thus, in the above experiment, the friction produced great heat in the glass, the string emitted a burnt odour and broke, and great expansion of the glass was produced. When the outer surface was suddenly cooled by the cold water, the expanded particles at once contracted, and more rapidly in the external particles than in those of the inner surface, causing the fracture of the glass, and the more easily the greater its thickness. If the temperature had been slowly reduced, it would not have broken.

Thus it is obvious (*a*) that glass and porcelain vessels intended for sustaining high temperatures, such as flasks, alembics, retorts, capsules, &c., should be thin, particularly at the bottom; and (*b*) that, when used, they should always be gradually heated and cooled.

The above method of heating glass by a piece of string

Fig. 12.



furnishes the chemist with a simple expedient for removing stoppers which are too firmly fixed in the bottles to be taken out by turning or tapping them. All that is necessary is to wind a piece of thick string round the neck of the bottle, and move it quickly until sufficient heat has been produced to loosen the stopper.

No two solids expand alike; the metals expand the most, and all solids less than fluids.

Expansion by Cold.—A remarkable exception to this law of expansion by heat and contraction by cold occurs in the case of water.

Experiment 4.—A large flask is arranged as directed in Experiment 2, page 28, but inserting a cylindrical thermometer, *a*, through a second hole made in the cork. The flask is filled with water to the top of the tube *b*, and placed in a vessel filled with ice. A strip of paper may be pasted on this tube, upon which the level of the water may

be marked as the thermometer falls. The water as it cools will sink in the tube until the mercury stands at 39°F. (4°C.); yet, on cooling still more, it does not fall any further, as we should expect it would, but, on the contrary, it begins to rise again, and continues to do so till it reaches the freezing point. At 32°F. (0°C.) it stands at the same point as when its temperature was at 46°F. (8°C.). Water is accordingly most dense at 39°F. (4°C.); all other liquids continue to increase in density as they cool.

However unimportant this exception may appear at first, our admiration must be excited when we reflect upon its consequences. Were it not for this, our country would have the climate of Greenland. The freezing of our waters, as the winter sets in, is principally owing to the coldness of the atmosphere. Consequently, the upper part of the water is colder and heavier, and sinks to the bottom; the warmer water ascends, becomes cold, and also sinks. If the water continually became denser, to its freezing point, this circulation would continue till the whole mass of water to its greatest depth reached freezing point, and a few cold days would suffice to convert all our ponds, lakes and rivers into ice. This does not happen, because the circulation ceases when its temperature has fallen to 39°F. (4°C.), when the water, though yet colder, becomes lighter, and floats on the surface. Thus, freezing can only take place at the surface, and the ice be but gradually formed. At a small depth below the ice the water generally retains the temperature of 39°F. (4°C.).

Expansion of Gases.—Experiment 5.—Dip a glass tube, provided with a bulb, into water, and heat the bulb gently; a part of the air is expelled, and escapes in bubbles through the water; consequently, there is not room enough in the bulb for the heated air; but it requires a larger space than it did in its cold condition. It follows from this, also, that the warm air is lighter than cold. If the lamp be removed, the air remaining in the bulb will contract on cooling, and water will be

Fig. 13.



pressed up into the bulb, replacing the air which has been expelled. Such an apparatus, when the tube is small, may be used as an *air thermometer*. When the bulb is heated, the liquid descends in the tube, and *vice versa*.

All gases expand alike when equally heated, differing in this respect from solids and liquids. The extent to which they expand is moreover much greater than with either of the other forms of matter. A gas at freezing point expands $\frac{1}{491}$ of its volume for every degree Fahrenheit ($\frac{1}{273}$ of its volume for every degree Centigrade) that is added to its temperature, so that it is very easy, knowing the volume that a certain weight of gas occupies at one temperature, to calculate what its volume would be at some given standard temperature. In experiments on gases their volume is always given (unless otherwise stated) at 0° C. or 32° F., which is therefore said to be the *standard of temperature*.

Example.—A certain weight of gas measures 23 cubic inches at 65° F.: what would be its volume at 32° F.?

491 volumes at 32° become 492 at 33° , 493 at 34° and so on, increasing 1 volume for every 1° F. Therefore, at 65° , the 491 volumes would become $491 + 33 = 524$ volumes. Hence the calculation is a mere rule of three sum:—

Vols. at 65° . Vols. at 32° .

$$524 : 491 :: 23 : x = 21.5 \text{ cubic inches.}$$

Similar calculations may of course be made in the Centigrade scale, as the following example will show.

We have 100 measures of a gas at 50° C.: what would be its volume at 0° C.?

273 vols. of gas at 0° C. = 274 vols. at 1° , 275 at 2° , and so on. Heated from 0° to 50° , 273 vols. would become 323 vols. Therefore—

$$323 : 273 :: 100 : x = 84.5 \text{ measures.}$$

We can in this way find the volume that a gas would occupy at any given temperature, although it most frequently happens that we want to find its volume at freezing point.

We have 100 volumes of gas at 17° C.: what would be its volume at 193° C.?

273 vols. at 0° C. = 290 vols. at 17° , and 466 vols. at 193° . Therefore—

$$290 : 466 :: 100 : x = 160.7 \text{ volumes.}$$

These calculations are generally spoken of as “corrections for temperature.”

Change of State as produced by Heat.

(1) *From Solid to Liquid.—Melting of Solids.*—Expansion is the first general effect of heat; but in many solid bodies another effect is observed; they change their state of aggregation, they become liquid, they *melt*. Many of them become *soft* before melting, so that they can be kneaded; for instance, butter, glass, and iron; in this condition, glass can be bent and moulded like wax, and iron can be forged.

Experiment 1.—Hold a piece of a small glass tube in the upper part of the flame of a spirit-lamp, revolving it slowly between the fingers; when red-hot, it will be so soft that it can be bent into any shape desired. Thus are easily formed any of the numerous bent tubes required in chemical experiments. For softening larger tubes, a flame with a current of air passing through it, or a Bunsen burner must be used, as either gives a much stronger heat than the simple lamp. To break a glass tube, a scratch is made upon it with a three-cornered file, at the place to be broken, and then it can be parted by gently bending it with both hands.

Most solid bodies become *suddenly fluid*, as ice, lead, &c.

Experiment 2.—Place one vessel containing snow or ice, and another containing a piece of a tallow-candle, on a warm stove, dipping a thermometer from time to time into the melting substances; the temperature will remain stationary in the first vessel at 32° F. (0° C.), in the other at about 100° F. (38° C.), so long as any ice or tallow remains unmelted, but when the melting is complete it will commence rising. The degree of heat at which a body melts is called its *melting point*. Every substance has its own melting point, sometimes above and sometimes below the freezing point of water; for example, lead melts at 620° F. (327° C.), silver at above 1800° F. (1000° C.),

solid mercury at -38° F. (-39° C.). If the two vessels containing the melted ice and the melted tallow are placed in the cold, it will be observed that the tallow soon hardens at about 95° F. (35° C.), but the water does not until the mercury has fallen to 32° F. (0° C.). Thus the congelation of fluids takes place at about the same temperature as that at which they pass from the solid to the fluid state.

Many substances, charcoal for instance, have never yet been melted, and others have never been frozen, as, for instance, alcohol; but it is very probable that, when some method of producing still greater degrees of cold and heat is discovered, we shall succeed in rendering all solid bodies liquid, and all liquids solid.

Latent Heat.—Experiment 3.—Put two vessels of equal size on the top of a warm oven, one of them containing a pound of ice at 32° F. (0° C.), and the other a pound of water at 32° F. (0° C.); when the snow is melted, remove them both. By the touch merely it will be perceived that the snow-water is still cold, while the water in the other vessel has become warm; and the thermometer will indicate that in the former the temperature is at 32° F. (0° C.), in the latter at 176° F. (80° C.). As both vessels have received equal degrees of heat, and when placed upon the oven were of the same temperature; the question suggests itself, What has become of the 144° of heat imparted to the vessel filled with ice? The reply is, This heat has been absorbed by the ice, thus converting it into a fluid,—melting it.

Experiment 4.—Put one pound of ice at 32° F. (0° C.) into the vessel containing the water heated at 176° F. (80° C.), and then examine it with the thermometer; as soon as all the ice has disappeared, the quicksilver will have fallen to the freezing point. Consequently, the ice has taken from the hot water 144° F. (80° C.) of heat, and has thus become liquid.

The heat which is lost in the melting of ice and other solids is employed in overcoming the cohesion of the solid. The force of cohesion is overcome in the manner described in the previous chapter, and when the liquid again becomes solid the heat is reproduced; the exertion of the

cohesive force being attended with the evolution of heat. Heat lost during liquefaction is often called *latent heat*, because it was formerly supposed to lie hidden, or stored up in the liquid. In a limited sense the term may still be retained, for although we now know that heat is nothing but motion, there can be no doubt that the molecules (particles) of the liquid are in incessant motion, which motion they lose to a great extent when the substance becomes solid.

(2) *From Liquid to Gas.—Boiling of Liquids.*—Most liquids, when heated to a certain temperature, different for every liquid, boil and become converted into vapour or gas. It must be understood that there is no real difference between the condition of a gas and of a true vapour, except in regard to the temperature at which they are formed. All the so-called gases can by great cold and pressure be converted into liquids; and, on the other hand, when a liquid is converted by heat into vapour, the vapour has all the properties of a true gas. In fact, some substances are known which are liquid in cold weather and gaseous in hot. Gases which become liquid at ordinary temperatures are generally called vapours.

The phenomena of boiling are most conveniently studied in the case of water.

Experiment 1.—Two-thirds fill a flask with spring water, and heat it gently over a lamp (Fig. 14). In a short time numerous little bubbles will appear on the walls of the flask, which will gradually increase in size, and rise towards the surface. These bubbles consist of the gases of air, which are expanded by the heat and expelled from the water. All spring-water contains gases in solution, and to these is chiefly due its refreshing taste, which is not found in boiled water or in that which has been standing for some time. Afterwards, when the water has become quite hot, larger bubbles appear on the hotter part of the flask, which also ascending, become smaller, and entirely disappear before reaching the surface of the water; they consist of gaseous water (steam), which con-

Fig. 14.



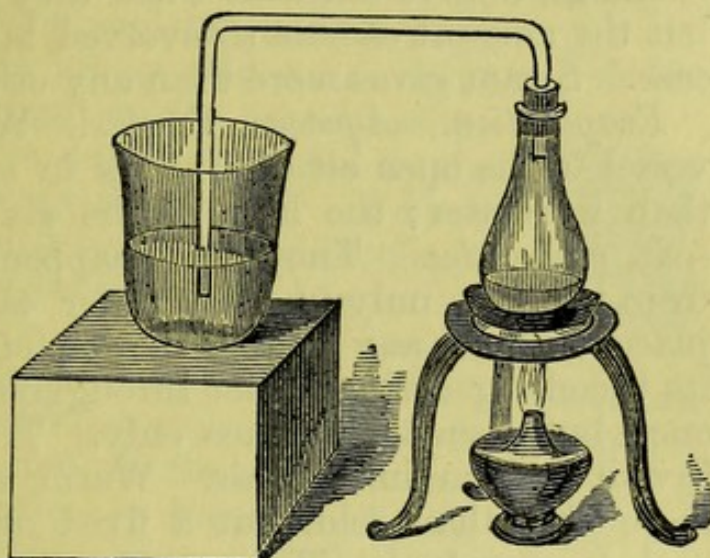
denses as it comes in contact with the cooler liquid above. The collapsing of the particles of water at the places where these steam-bubbles disappear occasions that peculiar noise which precedes boiling, and which is commonly called the *singing* of the water. When the whole mass of water is heated to 212° F. (100° C.), these bubbles no longer condense, but rise to the surface, where, surrounded by a thin film of water, they remain quiescent for a few seconds, and then, their watery mantle again sinking, they finally burst. This is the *boiling* of water. It boils at 212° F. (100° C.); other liquids boil more readily—alcohol, for instance, at 176° F. (80° C.); others again with greater difficulty—mercury, for instance, at 675° F. (357° C.).

Latent Heat of Steam.—The space above the boiling water in the interior of the flask appears empty, but it is in fact filled with aeriform water, which has displaced the air that was in it. This aeriform water is called *steam*. It is almost 1700 times lighter than water, because a measure of water yields nearly 1700 measures of steam at 212° F. (100° C.). Within the flask the steam is transparent and invisible, but in the open air it ascends in the form of white clouds, which greatly increase if cold air is blown into the flask by means of a glass tube; for on cooling the transparency of the vapour is disturbed, on account of the formation of drops of water, so small and light as to float in the air. Clouds also consist of this partly-condensed vapour. As the condensation increases, the drops become so large and heavy, that they descend as rain. A thermometer immersed in boiling water indicates 212° F. (100° C.); if placed in the steam immediately above, it shows the same; and this temperature will not rise higher, however long the boiling be continued, or however strongly the flame of the lamp be urged. This is similar to what occurs in the melting of snow; heat disappears, and its disappearance proceeds from the same cause in both cases; steam requires heat for its production, and this heat disappears, or becomes *latent* during its formation, and reappears, or becomes *sensible* during its condensation. We have, in fact, to study the latent heat of steam as well as the latent heat of water.

Experiment 2.—Adapt the shorter limb of a bent glass

tube, by means of a perforated cork, to the neck of a flask, and pass the longer limb to the bottom of a beaker-glass or common tumbler. Pour into each of these two

Fig. 15.



vessels five ounces of ice-cold water, and gradually heat the flask upon a tripod until it boils. Note the time required for this operation. Continue the process until the water in the beaker-glass begins to bubble, and note also the time, which will be found the same as that required for boiling the water in the flask. The steam formed in the flask has no other outlet than through the tube into the water, where it condenses, until the contents of the second glass reach the temperature of 212° F. (100° C.), and boil:

Both of the vessels must now be weighed; and it will be found that the flask weighs about one ounce less, and the beaker-glass about one ounce more, than before; consequently, one ounce has passed from the former as steam, and has been condensed again in the latter; and yet this ounce of steam has raised five ounces of water from 32° to 212° F. (0° to 100° C.), that is, through 180° F. (100° C.). Now five ounces raised 180° is the same as one ounce raised $180 \times 5 = 900^{\circ}$, so that the ounce of steam in condensing gave out 900° F. (500° C.) of heat. This is expressed by saying that the latent heat of steam is 900° F. (500° C.). More exact experiments give as the true number 965° F. (536° C.). The property of steam to give out a large quantity of heat during condensation peculiarly adapts it for the heating of other bodies, the burning of them being thus guarded against, as the heat of steam in open vessels can never exceed 212° F. (100° C.). Apothecaries avail themselves of steam in the preparation of infusions, decoctions, and extracts; it serves for many of the

processes of cookery, and for the distillation of spirits and other liquids; it is employed in dyeing and bleaching establishments, and is often resorted to for heating apartments, buildings, laundries, &c.

All gases give out heat when they condense into liquids, but the amount of heat so evolved is not known in every case. Steam gives more than any other known gas.

Evaporation.—Aqueous Vapour.—Water exposed in a vessel to the open air disappears in summer more rapidly than in winter; the heat of the air converts it into gas—it *evaporates*. The same happens as in evaporation over the fire, only in the former case evaporation takes place without any visible motion of the water, owing to its becoming aeriform, not throughout the whole mass at once, but upon the surface only. The vapour rises in an invisible form in the air. Warm air, indeed, takes up more of it than cold, but a fixed quantity of it only for each temperature. Thus one hundred measures of air at 32° F. (0° C.) absorb two-thirds of a measure of vapour; at 50° F. (10° C.), one measure and a quarter; at 68° F. (20° C.), two and an eighth measures, &c. If the air has not absorbed all the vapour which it can, it eagerly takes up more, as, for example, when one hundred measures of air at 68° F. (20° C.) contain only one or one and a half measures of vapour; it is then called *dry air*, and wet articles are soon dried in it by rapid evaporation. But if it be already saturated with vapour it is called *moist air*; and damp articles cannot be dried in it, or at least but slowly. If yet more vapour be added to this saturated atmosphere, or if it be cooled, then the excess separates in visible particles, called *mist* or fog when they lie near the surface of the earth, and *clouds* when they float in the higher regions of the atmosphere. The white smoke which in winter is seen rising from steampipes, the visibility of the breath in frosty weather, and the smoking of rivers in winter and after a storm, are phenomena of the same kind.

If the cooling of the air is occasioned by a cold solid body, the vapour is then condensed in small drops of water, as may be observed on the outside of a cold glass when brought into a warm room, and the deposit of moisture on the inside of our window-panes, when cooled by the ex-

ternal cold air. The temperature at which this occurs is called the *dew-point*, signifying the temperature at which the air is saturated with vapour.

Experiment 3.—Fill a tumbler one quarter full with cool water, place in it a thermometer, and at short intervals gradually add ice or cold water, until moisture begins to deposit on the outside of the glass. Then observe the degree indicated by the thermometer, which is the dew-point. If much cold water must be added before the glass clouds over, that is, if the dew-point is much lower than the temperature of the air, fair weather may be expected; while, on the contrary, if the difference between the dew-point and the temperature of the air be but slight, rain may soon be expected, as then the air requires but a slight addition of moisture or increase of cold to become saturated. Instruments by means of which the amount of moisture in the air is ascertained are called *hygrometers*. Many substances readily imbibe moisture from the air, and become damp; such bodies, for instance, as catgut, carbonate of potassium, sulphuric acid, fresh barley-sugar, &c., are called *hygroscopic*.

Fig. 16.



Evaporation may be accelerated, not only by heat, but also by a current of air, because by this means the air above the surface of the fluid, which is charged with vapour, is removed and replaced by a drier, and, as it were, more thirsty air, which takes up the vapour more rapidly and abundantly than the former. For this reason, the earth dries rapidly after rain, when followed by a high wind, and hence it is necessary in kilns, laundries, drying-rooms, &c., to arrange them in such a manner that the air, when saturated with moisture, may be constantly replaced by dry air.

That *heat disappears* during slow as well as rapid evaporation may be readily illustrated by the following experiment.

Experiment 4.—Fill a tube half full of water, and fasten securely round the bulb of it a piece of wadding; saturate the wadding with cold water, and then twirl the tube rapidly between the hands; presently the water in the tube will become sensibly colder, and the degree of cold may be accurately determined by the thermometer.

Moisten the wadding with ether, a very volatile liquid, and twirl it again in the same manner as before; by which

Fig. 17.

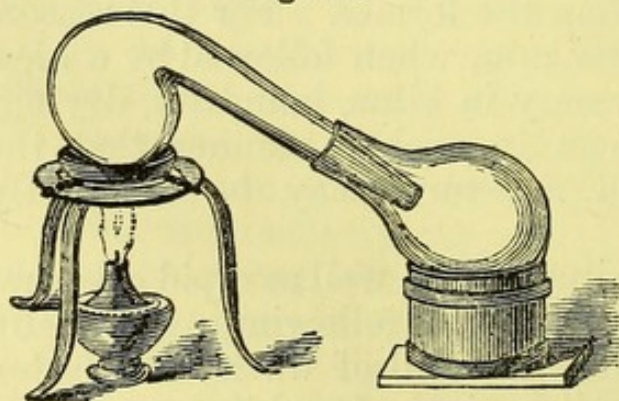


means its contents, even in summer, may be converted into ice. Water evaporates slowly, ether rapidly; and both require heat for their conversion into vapour, and in the above experiment they obtain this heat from the water in the bulb, which is of course the reason of the water becoming cold. On this principal, one feels cool on just leaving the bath, when invested in damp garments, or when the floor of a hot apartment is sprinkled with water. It explains, also, how man is enabled to support the scorching sun of the hottest climates, and even to endure a heat of 212° F. (100° C.), without his blood exceeding the temperature of from 100° F. (38° C.) to 104° F. (40° C.); it is owing to the more copious perspiration, which, by evaporation, renders all the heat above 104° F. (40° C.) latent. If we blow on hot soup, it is also the increased evaporation which cools it more rapidly; but if we blow on the cold hands in winter, they become moist and warm, because the latent heat contained in the vapour of the breath is set free, as the vapour is condensed into water.

Distillation.—If the condensation of watery vapour be carried on in a closed vessel, the water may be collected as it forms.

Experiment 5.—A small glass retort is half filled with

Fig. 18.



water, and heated; the steam, as it forms, passes through the neck of the retort into a glass receiver, contained in a vessel filled with cold water, and is there condensed. That the refrigeration may take place more rapidly, the receiver is covered with

coarse blotting paper which is frequently moistened with cold water. This operation is called *distillation*, and the pure water obtained is said to be *distilled*. It is purer than spring-water, for this reason, that the non-vo-

lative, earthy and saline portions contained in all spring-water do not evaporate, but remain in the retort. By this means also very volatile bodies can easily be separated from less volatile ones; as in the distillation of brandy, where the more volatile spirit is partially separated from the less volatile water. Copper stills are usually employed for distillation on a large scale, and for condensers vats are constructed, holding serpentine pipes, or *worms*, which present a greater condensing surface than if the pipe had passed straight through the vat. The cold water with which the vats must be filled is very soon warmed by the heat liberated in the condensation of the steam, and must be renewed by leading off the hot water from above, and letting in a fresh supply of cold water beneath.

DIFFUSION OF HEAT.

Conduction of Heat.—Experiment 1.—A test-tube, nearly filled with water, is held over a spirit-lamp, in such a manner as to direct

Fig. 19.



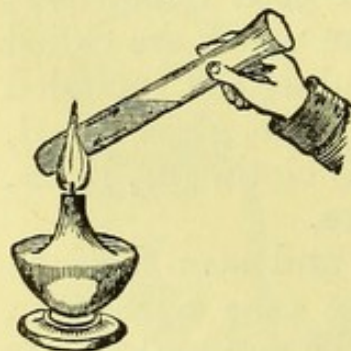
the flame against the upper layers of the water; the water will boil at the top, but remain quite cold below. If mercury is treated in a similar way, its lower layers will gradually become heated. The particles of mercury will communicate the heat to each other, but not so the particles of water. Substances through which, as in mercury, heat rapidly passes, are called *good conductors*; but bodies which comport themselves like water are called *bad conductors* of heat. In the former class are included principally the metals, and in the latter, stone, glass, wood, snow, water, and especially soft bodies, as cloth, fur, linen, straw, paper, ashes, &c., and also the gases.

The *good conductors* are readily heated, and soon become cold again, as is well known to be the case with iron stoves. A piece of iron feels hotter in the sun and colder in the shade than a piece of wood at the same temperature. The explanation of this delusion of the sense of touch, is that the warm iron conducts the heat more rapidly to the hand, while the cold iron withdraws it more rapidly than the wood is capable of doing.

The *bad conductors* of heat become only slowly heated, and also slowly cooled; for this reason, stoves constructed of brick (the Russian stove) and those made of Dutch tiles, a preparation of clay, retain their heat longer than iron stoves. Bad conductors are frequently employed both for preventing the quick heating and the quick cooling of bodies. Vessels of glass and porcelain are placed on sand (a sand bath) or asbestos, to heat them gradually, and thus guard against their breaking. If a hot liquid is to be poured into them, it must be done by small portions at a time, twirling the vessels round for some minutes before adding more.

On removing vessels from the fire, the precaution should be taken never to place them while hot on metal or stone, but always on some bad conductor, such as straw, wood, paper, cloth, &c.; as they are often cracked by sudden cooling and contraction, which is frequently produced even by a current of cold air. Doors of furnaces, ladles, &c., are provided with wooden handles, to prevent those using them from being burnt. Should a person desire to hold a flask or a test-tube while liquids are boiling in them, he must wrap round them several folds of paper, or tie round them a piece of twine, in order that they may serve as a bad conductor between the glass and his fingers. By inclosing substances in bad conductors, the entrance of cold, or, more correctly, the departure of heat, may be prevented; this principle is illustrated in our method of

Fig. 20.



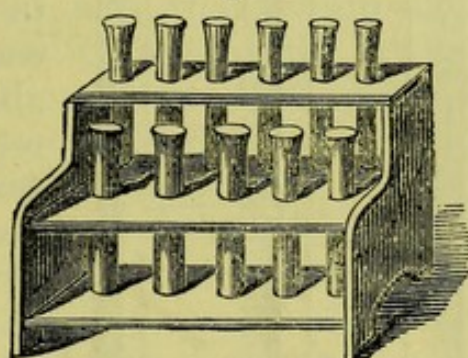
clothing, in the protection given to our wells and trees by covering them with straw, in the preservation of plants by snow, and in numerous other phenomena of daily occurrence. Hence bad conductors are frequently called preservers of heat.

Convection of Heat.—Liquids and gases, though very bad conductors, readily become heated by a process called *convection*, which is a necessary consequence of the expansion produced in them by heat.

Experiment 2.—Water, to which some sawdust has been added, is heated in a test-tube over a spirit-lamp. The tube is held by the upper part, and rotated for some

minutes between the fingers, that the flame may have equal access to all the lower parts of the tube. If the water be carefully observed, it will be seen that the saw-dust ascends on the upper surface of the liquid, and descends in the lower strata; the warm water expands, and, becoming lighter, rises upwards, while the colder, and consequently heavier, water sinks; the water *circulates*. In consequence of this circulation, the heating of fluids takes place more rapidly when the heat is applied *beneath*. If the *upper part* of a test-tube full of water is heated by the lamp, the water will boil at the top while it remains quite cold, and may even

Fig. 21.



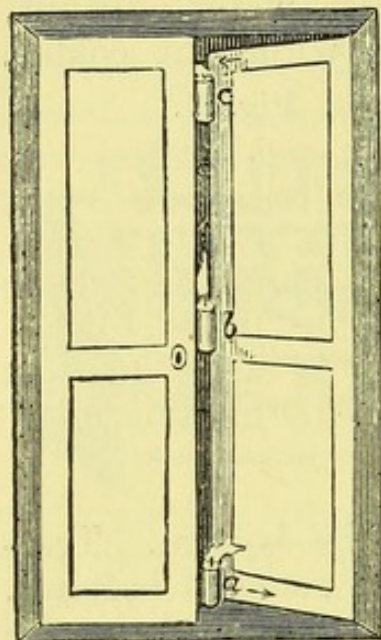
contain ice, held down by a wire, at the bottom. Test-tubes are cylindrical glass vessels with rounded bottoms. To prevent their breaking on the application of heat, the bottom must be thin, and blown of a uniform shape. A simple wooden rack, as in the annexed figure, serves as a convenient stand for them.

Convection in Gases.—Currents of Air.—A great many phenomena of daily occurrence may be explained by the difference in lightness between warm and cold air. When a fire is lighted in a stove for the heating of an apartment, the air immediately in contact with the stove is first heated, becomes lighter, and ascends; colder air rushes in to supply its place, and this likewise becomes heated and ascends; consequently a constant circulation of air is kept up. By a similar circulation, the whole atmosphere of the earth is kept in continual motion. At the equator, the strongly heated air ascends and moves in the upper regions of the atmosphere towards the poles, while in the lower regions the current of cold air flows from the arctic zone towards the equator, in order here to restore again the equilibrium, disturbed every moment by the ascent of the warm air. These regular currents of air, the direction of which is somewhat diverted by the revolution of the earth on its axis, are called trade-winds.

In every heated apartment, a difference between the heat

of the air near the ceiling and that near the floor is very perceptible. If a door or window in such a room be opened, a current of air is produced, the direction of which may

Fig. 22.



easily be perceived by holding a lighted candle in the opening; the flame, when held above, at *c* (Fig. 22), is blown *from* the room; when placed below at *a*, it is blown *into* it; consequently, the light warm air above rushes out of the room, and is replaced by heavier and colder air from below. A draught of air is also noticed in passing from the sunshine into the shade; where the sun shines, the warmer air ascends, and the colder air from the shade supplies its place. For the same reason, a current of air is produced wherever a fire is burning, in every stove, and round every lamp. To *ventilate* a room properly, the hot and foul air should be allowed to escape from some aperture near the ceiling, while cool and fresh air is admitted near the ground.

The air-balloons, first constructed by Montgolfier, strikingly show how buoyant air may be rendered by heat; these are caused to ascend merely by filling them with air, kept continually hot by a fire beneath.

Radiation of Heat.—By conduction, bodies can communicate or abstract heat only when in contact. But heat is felt even at some distance from a fire or from a heated stove, and the earth is warmed by the sun, although a space of millions of miles is between them. This sort of heating is called *radiation of heat*.

Experiment 3.—Envelop three tumblers with paper, one with silver paper, another with white, and a third with dull black paper, and place them in the sun; a thermometer will indicate that the tumbler with the black paper is heated the most, and that with the silver paper the least, and yet all these vessels have been equally exposed to the rays of the sun. This difference is explained on the principle, that some substances have the power of absorb-

ing the rays of heat, while others throw them off again, or *reflect* them. Good reflectors of heat, such as bright metals, are the worst absorbers, for a thing cannot absorb the heat and also reject it.

If hot water is poured into the tumblers enveloped with paper, and the cooling of it noted by the thermometer, the contrary effect will be observed; the glass covered with black paper will first become cold, and that wrapped in silver paper the last; because bodies with dull surfaces radiate the heat more rapidly than those with polished surfaces. Good absorbers of heat are therefore seen to be good radiators, and good reflectors bad radiators. For this reason, coffee retains heat longer in a bright than in a tarnished pot; a stove of glazed Dutch tiles remains hot longer than another of unglazed tiles; a smooth sheet-iron stove, longer than a similar one of rough cast-iron, &c.

The radiation of heat enables us to explain some of those common natural phenomena which otherwise would remain obscure. Why do not the rays of the sun, even in the hottest summers, melt the snow upon the tops of high mountains, which are nearer than the level portions of the earth to the sun? Because they only heat those bodies which can absorb their warmth, as the rough surface of the earth. The snow is indeed struck by the rays of the sun, but being a white and shining body it reflects them and remains cold.

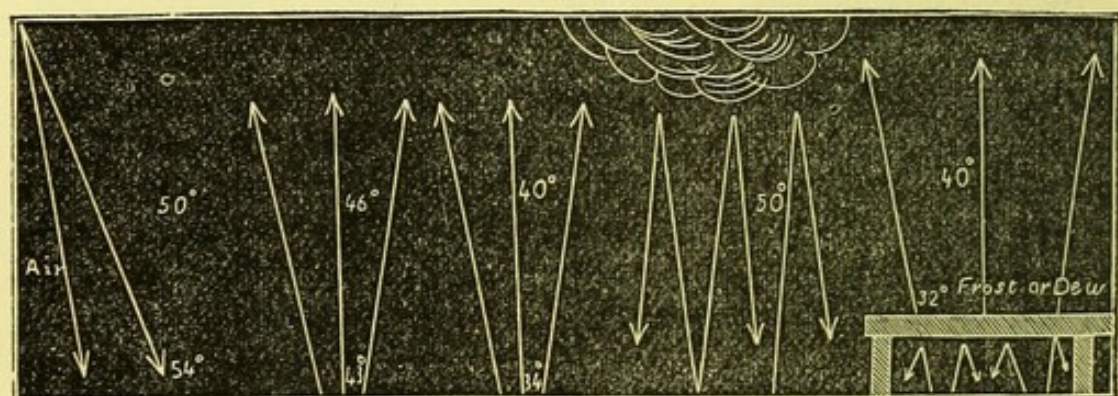
Formation of Dew.—When the surface of the earth has become warm, the air is heated by it; hence, during the day the lower strata will always be warmer than the upper. But a change takes place after the sun has gone down. The earth continues to radiate heat without receiving any in exchange, and its temperature consequently diminishes. On the other hand, the air does not so readily part with its heat, and therefore it retains during the night a higher temperature than the surface of the earth; it is only cooled where it comes in contact with the colder earth. If this cooling should reach the dew-point of the air, then the vapours are condensed, on the soil or on the plants growing upon it, in the form of small drops, just as a tumbler is covered with vapour when brought from a cold into a warm room—dew forms. If

the temperature of the earth sinks in the night to the freezing point, or below it, the aqueous vapour is deposited in a solid form, and is called *hoar-frost*.

The radiation of heat from the earth is greatest when the weather is clear and serene, but it is obstructed by clouds and wind. Hence the most copious deposit of dew takes place only in clear and quiet nights. The clouds serve as a screen in reflecting back the rays of heat to the earth, so that it is but slightly cooled. The same effect is produced by the mats, straw, and boards with which the gardener covers his young plants, to protect them from the late frosts of spring, or from freezing. The annexed figure, in which arrows denote the direction of the rays of heat, will serve to render this process more intelligible.

Sunbeams.

Fig. 23.



59° Surface of the earth.	41° Dew.	32° Frost.	54° No dew or frost.	41° No dew or frost.
In the day-time.	In clear and serene nights.	Cloudy or windy nights.	Clear nights.	Soil protected.

SOLUTION AND CRYSTALLIZATION.

These subjects are so closely connected with that of heat that they may conveniently be studied in this place.

Solution.—Water can dissolve many bodies, and unite intimately with them, without losing its transparency. Such combinations are called solutions. If rain-water meets with soluble substances, either in the earth or in the rocks through which it oozes, it dissolves them; and this explains why all spring-water, on evaporation, yields an

earthy or saline residue. Frequently this residue, particularly when it contains lime, is so altered during evaporation, that it can no longer be dissolved in water, and forms a hard incrustation round the inner sides of the vessels used in cookery. The springs of Carlsbad deposit so much residue, that articles immersed in them appear in a short time to be externally petrified or incrustated. If water is unusually rich in soluble substances, especially such as possess medicinal properties, as, for example, iron, sulphur, &c., it is called *mineral water*, and the springs from which it issues are called mineral springs. A pound of sea water contains about half an ounce of substances in solution.

Experiment 1.—Pour a teaspoonful of slaked lime into a bottle, and fill it with water; cork it up, and, after shaking it for some minutes, let it stand until the water has become perfectly clear. By carefully inclining the bottle, most of the liquid may be poured off free from the sediment. This operation is called *decantation*, and the clear liquid is *lime-water*. Lime is but slightly soluble, about 600 parts of cold water being required to dissolve one part of it; the excess remains undissolved, and as lime is heavier than water, it settles at the bottom. That a portion of it has been dissolved is known by the peculiar taste imparted to the liquid. This taste is called *alkaline*.

Keep a part of the lime-water in a well-stopped bottle for future use; it will remain transparent and clear. Pour the remainder into a tumbler, and expose it to the air; the water soon becomes turbid, and covered with a film, which gradually grows thicker, and settles at the bottom. When, after some days, the water has become clear again, it will have lost its alkaline taste; the lime dissolved in it, having been chemically changed by the air and rendered insoluble, will be found as a powder at the bottom of the tumbler.

Experiment 2.—Put into a flask half an ounce of litmus; pour over it three ounces of water, and let it remain in a warm place until the liquid has obtained a dark-blue colour. Litmus consists of a blue colouring-matter, which is soluble in water, and is hence taken up by it; it also contains some earthy matter which is insoluble, and is

deposited as a slimy mass. These two substances might be separated from each other, as in the former experiment, by decantation, but it can be done more readily by *filtration*. For this purpose cut a piece of blotting-paper into

Fig. 24.

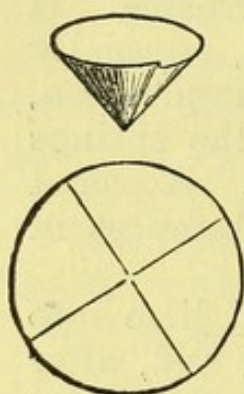


Fig. 25.



a circular shape, fold it together twice, and then place this filter into a glass funnel. That the paper and the glass may not come into too close contact, place between them thin pieces of wood or glass; a piece of string must also be inserted between the funnel and the neck of the flask into which the liquid is to be filtered, to allow an opening for the escape of the air from the flask, as otherwise the

fluid could not flow in. The filter, which must never be higher than the top of the funnel, is first moistened with water before the fluid is poured upon it. Blotting-paper consists of fine linen or cotton-fibres matted together, between which are small interstices or pores, through which liquids, but no fine solid particles, can pass; these remain on the filter. Writing-paper cannot be used for filtration, as its pores are filled up by size or starch.

Experiment 3.—Pour a part of the solution obtained into a saucer, and pass strips of fine blotting- or of letter-paper one or more times through it, until they have acquired a distinct blue colour. Preserve these strips, after being dried in a box; they are called blue litmus, or test-paper; they are reddened by vinegar, lemon-juice, and all acid fluids, and serve to test a liquid, to ascertain whether it is *acid* (has an acid reaction).

Experiment 4.—Mix cautiously another portion of the solution with lemon-juice or vinegar, until the blue colour appears distinctly red; this also serves to colour paper. The red test-paper is used for the purpose of recognising a class of substances opposed to acids—that is, alkaline or basic bodies; these restore the original blue colour of the paper, as can be seen by bringing it into contact with lime-water or solution of caustic potash.

Experiment 5.—Add gradually, with constant agitation,

to one ounce of cold water, powdered saltpetre (potassium nitrate), as long as it continues to be dissolved—perhaps about a quarter of an ounce: if more is added than is necessary, it will remain undissolved at the bottom of the vessel. This solution is called a *cold saturated* solution. If this mixture be boiled, and saltpetre again be added, then about three ounces more will be required to saturate the water. A thermometer held in this *boiling saturated* solution will indicate about 240° F. (115.5° C.), while simple boiling water indicates only 212° F. (100° C.). *All saline solutions boil and freeze with more difficulty than water.* All bodies soluble in water behave in a similar manner—that is, they are soluble in it only in definite quantities, and in most cases hot water dissolves more of them than cold.

Experiment 6.—If the solution obtained in the last experiment be poured into a porcelain dish and be suffered to remain quiet until cold, then the three ounces of saltpetre which were last added separate again, not as a powder, but as regularly formed prisms. These prisms are six-sided, and are surmounted by faces similar to a roof; they are called *crystals of saltpetre* (Fig. 26.). All crystals are characterized by having planes, edges and angles, constructed, as it were, of simple triangular or quadrangular pieces, with polished surfaces. This symmetry is found even in the interior of them, as can easily be seen by holding a transparent crystal towards the light, and turning it slowly round; or breaking it, when the fragments will often exhibit the same regular form which characterized the whole crystal. Thus, in inanimate nature, a mysterious power exists, similar to that which compels the bees to construct six-cornered cells, and the potato to produce its five-angled corolla and five stamens, and by which the smallest particles of bodies, called *molecules*, are forced to arrange themselves in a fixed order, assuming a regular shape. But this can in general only be accomplished by a body in its fluid or aeriform state, since *free motion* of the molecules is essential. Time also is required for this operation; hence crystals are always more regular the more slowly they are formed. Many of

Fig. 26.



the splendid crystals which are dug from the depths of the earth, were, perhaps, very many years in forming.

Experiment 7.—Evaporate the liquid remaining above the crystals in the former experiment at a gentle heat, until a scum is formed on the surface, then remove it from the fire and let the solution cool, *stirring it constantly* with a wooden stick. In this way, instead of crystals, a *powder of saltpetre* will be obtained consisting of minute crystals.

The liquid just alluded to may be regarded as a cold saturated solution, containing about a quarter of an ounce of saltpetre. If, by evaporation only, so much water is left as is sufficient when hot to keep in solution but a quarter of an ounce of the saltpetre, then crystals begin to appear in the form of a film on the cooler surface, indicating the saturation of the liquid. If this again is allowed to cool quietly, a second crop of crystals will be obtained; but, by continual stirring, they are broken at the moment of their formation—by slow movement into a coarse, and by rapid movement into a fine powder. This may be called *interrupted crystallization*. Sugar presents a similar example; the same syrup, which, if cooled quietly, yields sugar-candy, when stirred, yields common loaf-sugar.

Experiment 8.—Put into boiling water as much common salt as will dissolve, and let the solution cool; very few crystals will form, because salt is nearly as soluble in cold as in hot water. Now evaporate one half of the solution over a spirit-lamp, and set aside the other half in a warm place; in the first case, mere irregular grains of salt will be obtained, but in the latter case, after some days, regular cubes of salt will be deposited.

Experiment 9.—Dissolve a spoonful of salt and one of saltpetre in lukewarm water, and put the solution in a warm place, that the water may gradually evaporate; the two salts, which are intimately mixed in the solution, will upon crystallization separate completely from each other. The saltpetre separates into long prisms, containing no trace of the common salt, and the latter separates into cubes, entirely free from saltpetre. Thus the particles of salt and saltpetre did not attract each other; but, upon

crystallizing out of the solution, the *homogeneous* salts assumed separately a regular form, precisely as if one only of these two substances had been dissolved.

PRESSURE.

The earth is surrounded by air, as by a mantle; this mantle is called the *atmosphere*, and extends to an indefinite height above the solid earth. The air possesses no colour, and is transparent; hence it is invisible, and its particles are so easily movable upon each other that it cannot be grasped by the hand. But it is rendered obvious that the air is material, and fills every space commonly called empty, by wrapping strips of moistened paper round a funnel, so that it may fit exactly into the mouth of a bottle; if the funnel be now filled with water, the fluid will *not* run into the bottle, as the air contained in the latter will not let it enter; but if the funnel be raised a little, the air escapes, and the water immediately rushes into the bottle. We learn also by the balance that a vessel apparently empty, *i.e.*, containing atmospheric air, weighs more than one which is really empty, as when the air has been exhausted from it. But air is so light that 800 measures of it weigh rather less than one measure of water, yet the atmosphere presses with great weight on the earth, and upon everything thereon. But this pressure is only noticed when the air is removed from a place, thus leaving it without counter-pressure. The total weight of the atmosphere has been estimated to be equal to that of a globe of solid lead sixty miles in diameter,

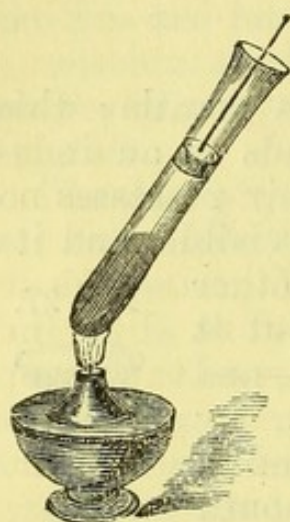
Fig. 27.



When a solid is immersed in a fluid (either a liquid or gas), it is pressed upon equally on all sides by the fluid with a force which depends partly upon the weight of the fluid and partly upon the depth to which the solid is immersed. A stone suspended one foot from the surface of water is only pressed upon with half the force that it would experience if sunk two feet; a stone sunk one foot in mercury would receive $13\frac{1}{2}$ times as much pressure as

it would if sunk one foot in water, because mercury is $13\frac{1}{2}$ times as heavy as water. The same thing is true in regard to air and other gases, and bodies near the surface

Fig. 28.

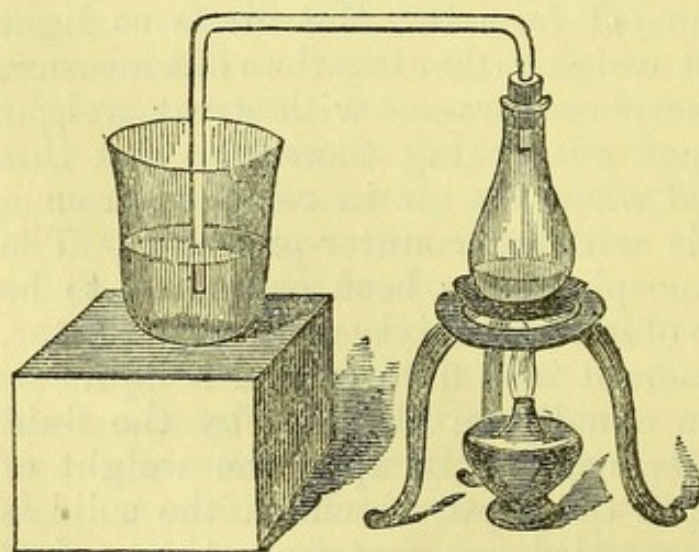


of the earth are pressed upon equally on all sides by the whole forty-five miles of air above,

Pressure of the Atmosphere.—Experiment 1.—Wrap some tow round one end of a stick and grease it with tallow, thus forming a plug, which must fit tightly into a strong test-tube. Boil some water in the test-tube, and when the air has been expelled by the steam, insert the plug; as the water cools the steam condenses, a vacuum is produced, and the plug will be pressed down upon the surface of the water; by heating, it is again forced up by

the steam thus generated, and by immersing in cold water it is again forced down. In consequence of the cooling and condensation of the steam a vacuum is formed, and, therefore, the counter-pressure against the weight of the exterior air is removed; the pressure of the latter, accordingly, forces down the plug. On this principle the

Fig. 29.



piston is forced up and down in the cylinder of many steam-engines.

This one-sided pressure often causes the ascent and reflux of liquids in tubes.

Experiment 2.—If water is boiled as was directed at page 39, by means of steam, and during the boiling the lamp

is removed, then the pressure of the air acting on the surface of the water in the beaker-glass will very soon force the water contained in it through the tube back

into the flask, which in a short time becomes quite filled with water. The counter-pressure of the steam must naturally decrease as it cools and condenses. As long as the lamp is under the flask, the pressure of the steam is stronger than that of the air, and the steam, being continually generated, forces the air previously contained in the flask into the water of the beaker-glass. This reflux of liquids is particularly to be feared, when such kinds of gases are conducted into water as are absorbed by it readily, and in large quantities. This is prevented by passing through the cork a second glass tube open at both ends, and letting it reach nearly to the bottom of the flask, by which tube air can penetrate into the flask as the pressure of the steam diminishes. This contrivance is called a *safety-tube*.

Fig. 30.



Barometer.—It can be proved by experiment that the atmosphere presses upon the earth with a force equal to that of a layer of mercury 30 inches deep, or a layer of water $13\frac{1}{2}$ times deeper (34 feet), water being $13\frac{1}{2}$ times lighter than mercury. The instrument by which the amount of atmospheric pressure can be observed and measured is called the *barometer*. Fill a glass tube, 32

Fig. 31.

inches in length, one end of which is closed, with mercury; close it with the finger, and invert it into a vessel of mercury; on removing the finger, the mercury will not run out, but will fall some inches, perhaps to *s* (Fig. 32). The height of the column of mer-

Air.
Water, 34 feet high.
Mercury, 30 inches high.

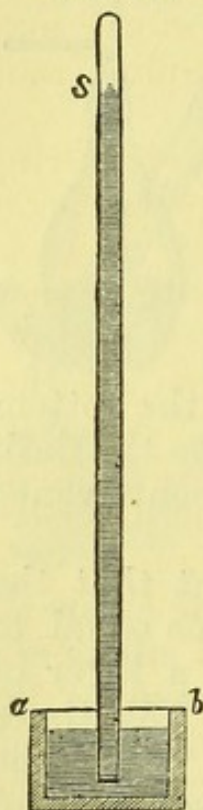


Surface of the earth.

cury, from *a b* to *s*, amounts to about 30 inches. The mercury does not fall lower, on account of the external pressure of the atmosphere, which is exerted on the mercury at *a b*, and not at *s*, since this end is closed. The column of mercury in the tube may be regarded as the counterpoise to the atmospheric pressure, and it is hence concluded that the latter exerts just as much pressure upon the earth as a column of mercury 30 inches high. If the tube be opened at the top, the

pressure of the air on both extremities being then made equal, the mercury will flow from the tube. The space above the mercury, at *s*, is a vacuum, and is called the

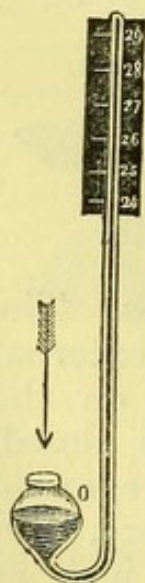
Fig. 32.



Torricellian vacuum, because it was first observed by Torricelli. In some barometers the tube is curved at the bottom, and provided with a bulb. This bulb is open at the top, and supplies the place of the vessel filled with mercury in the preceding figure. Here also the pressure is only exerted at one end, for the atmosphere can only press on the mercury contained in the bulb. The height from *o* (Fig. 33) to the top of the mercury amounts to about 30 inches.

If weights be placed in one pan of a balance, the opposite one will rise, but on their removal it will sink. The same thing happens with the barometer. Any increase in the weight or density of the air presses the mercury up, and the barometer rises; but any diminution of weight will make it fall. The height of the mercury may be read off by fixing to the upper part of the tube a scale divided into inches and tenths of an inch. The mean state of the barometer is at 30 inches, equal to 760

Fig. 33.



millimetres, and 31 is called a very high, and 29 a very low, state of the barometer. In this country, as a general rule, the north and east winds cause the barometer to rise, and the south and west winds cause it to fall. The former winds, blowing chiefly from the land, are cooler, and at the same time drier, than the latter, which pass over the ocean, there becoming saturated with moisture; the former likewise come from colder into warmer, while the latter, on the contrary, proceed from warmer into colder regions; by which the capacity of saturation for vapour is increased in one case and diminished in the other. Hence it is natural that, when north and east winds prevail, it should rain less frequently than during south and west winds; and that the former winds are dry, while the latter are damp.

Why water does not flow from a jar inverted over the pneumatic trough, why it continues to flow through a syphon after the air has been exhausted, why liquids will not run into a vessel when the air is confined, or why water will only rise to the height of 34 feet in a suction pump, are questions that scarcely require further explanation.

Elasticity of Gases.—All gases are perfectly elastic; they all contract equally under an increase of pressure, and expand when it is diminished. When a gas is collected in a bottle or other vessel the weight of it that the bottle will hold depends upon the pressure of the air at the time, which, as we have seen, is measured by the barometer. It is therefore of the utmost importance in measuring a gas to take account of the state of the barometer at the time. One hundred grains of air, for example, will occupy 333 cubic inches when the barometer is at 29, 322 when it is at 30, and only 311 when it is at 31 inches. As the alterations in volume are exactly in inverse proportion to the alterations of pressure, it is easy, after noting the volume of the gas and observing the height of the mercury in the barometer, to calculate what the volume would be at some standard pressure. The standard pressure at which gases are always measured is represented by 30 inches (or 760 millimetres) of mercury. The calculation required for this purpose is called the "correction for pressure." It must be applied, in addition to the correction for temperature (page 34), whenever gas is to be measured. It can always be made by the formula $V P = V^1 P^1$, in which V is observed volume, and P observed pressure. Hence

$$\frac{\text{Vol. at observed pressure} \times \text{Observed pressure}}{\text{Standard pressure}} = \left\{ \begin{array}{l} \text{Volume of gas at} \\ \text{standard pressure.} \end{array} \right.$$

For example, we have 333 cubic inches of air when the barometer stands at 29 inches. How many cubic inches would there be if the barometer stood at 30 inches?

Here we must multiply by the observed pressure 29, and divide by the standard pressure 30:

$$\frac{333 \times 29}{30} = 322 \text{ cubic inches (nearly).}$$

The corrections for pressure and temperature may be made separately, in either order, or may be combined into one, as in the following example:—

100 c.c. (cubic centimetres) of gas at pressure 700 m.m. and temperature 10° C. What would be volume at standard pressure and temperature?

For pressure,—

$$V^1 = \frac{100 \times 700}{760}.$$

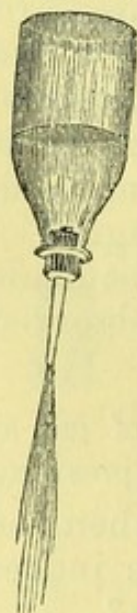
For pressure and temperature,—

$$\frac{100 \times 700 \times 273}{760 \times 283} = 89 \text{ (nearly).}$$

The ordinary *air pump* depends for its action in producing a vacuum upon the elasticity and pressure of the air.

If the pressure or tension of a confined quantity of air be increased, by compressing it either directly or by the addition of more air, it can be forced to stream out from a small opening with great rapidity, as is shown on a small scale in the common bellows, and on a larger scale in the blacksmith's bellows. Should there be water before this opening, the air will press it out in a jet or stream.

Fig. 34.



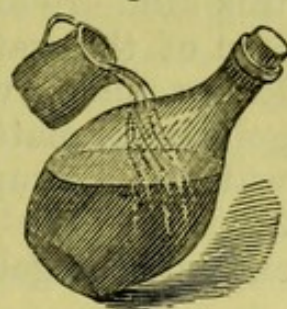
Experiment 3.—Take a piece of a fine glass tube, draw out to a point, and adapt it, by means of a perforated cork, to a bottle. Fill the bottle half full of water, and blow into it through the point of the tube; when the blowing ceases, the air will escape in a stream. But if the bottle be inverted as soon as the air is blown in, then the water will be spurted out by the compressed air above. Such an apparatus (the washing-bottle) was at one time employed for washing residues or precipitates remaining on filters, in order to free them from soluble matter. The washing-bottles now used have besides the jet, which is curved, a second tube by which air can be blown in and the water thereby forced out.

The pressure of the atmosphere exerts a great influence

on the boiling of water, and of other liquids. If water is made to boil when the mercury in the barometer is very low (in foul weather), brisk ebullition will take place at about 210° F. (99° C.); when the mercury is very high (in clear weather), boiling will not occur under 214° F. (101° C.).

Experiment 4.—Heat a flask half filled with water till the water boils briskly; then remove it from the fire and quickly cork it; the boiling immediately ceases, but will commence again if cold water be poured over the upper part of the flask. In this manner it can be made to bubble or boil, even though it be only lukewarm. There is no air in the flask, it having been expelled by the steam, and it could not re-enter it, on the cooling and condensation of the steam, on account of its having been closed. Consequently there is no pressure of air on the water, and it will boil even at a temperature of 68° F. (20° C.). The boiling ceased on account of the pressure of the steam upon the water; but the steam being condensed by the cold water, the pressure was so much diminished, that a portion of water again became aeriform with a boiling motion. In many manufactories, an appropriate apparatus (vacuum-pan) has been contrived for boiling and evaporating in a vacuum, as, for instance, in sugar-houses.

Fig. 35.



The air is densest at the level of the sea, and thinner in proportion to its distance from the earth, as there is less air above it. Hence the mercury will stand lower, and water boil more easily, on the top of a mountain than in the valley below. On the top of Mont Blanc mercury rises only to the height of 16 inches in the barometer, and water boils at 183° F. (84° C.). Hence the barometer and the boiling point of water may be employed for calculating the heights of mountains.

As water boils more easily under diminished pressure, so it boils with more difficulty when the pressure is increased. An increase of pressure can be produced, not only by the air, but by the steam of the water itself, if new steam be constantly generated, while the escape of that already

formed is prevented. This is best done by heating water confined in a strong and firmly closed vessel. For this purpose a Papin's Digester may be used, in which water may be heated to the temperature of 392° F. (200° C.), and, indeed, still higher, whilst in open vessels it cannot be heated above 212° F. (100° C.). If the amount of steam in it is twice as much as in an uncovered vessel, the pressure is said to amount to two atmospheres; if there is 3, 4, 5, 10, 20 times the quantity, there is said to be a pressure of 3, 4, 5, 10, 20 atmospheres. Vessels of this kind are often employed to effect a complete penetration of the water into solid and hard substances. Thus, for example, water at 212° F. (100° C.) dissolves the gelatinous matter only on the surface of the bones, whilst water at a temperature ranging from 230° F. (110° C.) to 248° F. (120° C.) entirely penetrates the bones, and extracts the gelatine also from the interior of them.

CHAPTER III.

FUNDAMENTAL LAWS OF CHEMISTRY.

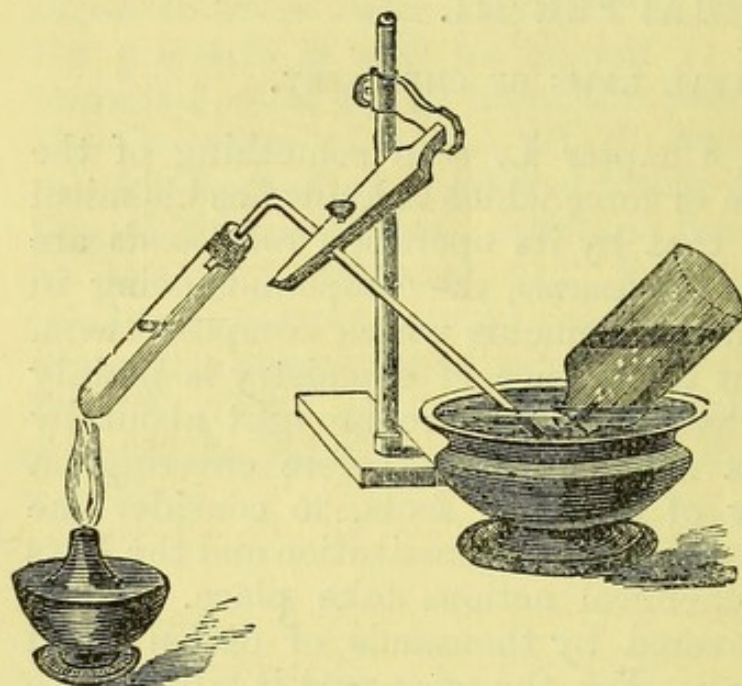
WE have already, in Chapter I., seen something of the operations of that form of force which is known as chemical force. We have seen that by its operation *compounds* are produced by the union of *elements*, the compounds being in all cases different from the elements which compose them. And we have seen that the science of chemistry is mainly concerned with the changes which are brought about by chemical force. It is now necessary, before entering on the systematic study of chemical facts, to consider the *theories* of chemistry—the laws of constitution and the laws according to which chemical actions take place. These laws have been discovered by thousands of difficult and laborious investigations. For the most part it will not be possible for the student to prove them to himself experimentally: the processes necessary for this are too difficult, and require apparatus of too expensive a kind. He must be content in the outset of his study to take the laws upon trust.

MODES IN WHICH CHEMICAL ACTION TAKES PLACE.

Combination.—*Experiment 1.*—The simplest way in which chemical action can take place is when two substances, elements or compounds, combine directly together and produce a single new compound. Examples of this have already been given at page 13, but another may be added here. Take a globule of mercury about the size of a very small pea, heat it in a test-tube until it boils, and then throw in about an equal weight of iodine. Shake the tube slightly, so as to bring the two into contact. They will combine, and a beautiful scarlet compound called mercuric iodide will be found on the sides of the glass.

Decomposition.—*Experiment 2.*—Another kind of chemical change is called *decomposition*. It consists in a separation, either partial or complete, of the constituents of a compound from one another. The following experiment will illustrate it. Introduce into a somewhat tall, but not too thin, test-

Fig. 36.



tube of hard glass, 108 grains of mercuric oxide. One end of a bent glass tube is adapted to it by means of a perforated cork, and the other end is conducted into a vessel filled with water. Either suspend the tube by means of a piece of string or wire, or support it by a retort-holder. A retort-holder is a wooden stand, pro-

vided with a movable vice, by which glass vessels can be held in the most convenient manner, as shown in the annexed figure. Then heat the test-tube carefully until all the mercuric oxide has disappeared. The red powder becomes black as the heat increases, and bubbles of gas escape, which are collected in a pint bottle held over the end of the tube, this bottle having been previously filled with water and then inverted into the bowl, after closing the mouth of it with the finger or a glass plate. No water will escape until bubbles of gas from the tube are passed into it, which, on account of their greater lightness, ascend and displace the water. When the evolution of gas ceases, cork the bottle and remove it. The first bubbles that pass over consist of the atmospheric air contained in the test-tube, and expanded by the heat, but *oxygen gas* quickly succeeds. This is *one* of the component parts of mercuric oxide, and can easily be recognized by the vivid combustion in it of a glowing

shaving. At the same time there is formed on the upper part of the test-tube a brilliant metallic mirror, which consists of mercury, the *second* element of the red oxide. When the latter has entirely disappeared, *immediately* withdraw the tube from the water, let the test-tube cool, and unite the mercury adhering to its sides into a single globule, by means of a feather. It should amount in weight to 100 grains; this, subtracted from the original weight, 108 grains, leaves 8 grains, the amount of the oxygen. The red powder consists of a brilliant heavy metal and of a gas, two entirely dissimilar bodies, and it has been *decomposed* by the heat. If these are chemically combined together by proper means, they will unite again and form the red oxide, a body in which the peculiar properties of the mercury as well as of the oxygen are entirely lost.

Double Decomposition.—Experiment 3.—It very often happens that two compounds are capable of decomposing one another, thus producing two new compounds. This kind of chemical action is called *double decomposition*. Weigh out 27 grains of mercuric chloride (corrosive sublimate) and 33 of potassium iodide. Dissolve them separately in water and then mix the solutions. A yellow precipitate will appear, which will quickly change to a beautiful scarlet powder, which will soon subside and leave the liquid colourless. The red powder is *mercuric iodide*, the same compound which was formed in Experiment 1. The clear liquid on evaporation will yield cubical crystals of *potassium chloride*. Mercuric chloride consists of the elements mercury and chlorine; potassium iodide, of potassium and iodine. When mixed, the mercury of the one unites with the iodine of the other, and the chlorine of the one with the potassium of the other.

Substitution, or Displacement.—Experiment 4.—Take a cold saturated solution (see page 51) of mercuric chloride, and immerse in it a strip of bright copper. Very soon the surface of the copper will become covered with a film of mercury, and in the course of a few hours the whole of the mercury will be separated in this way, and some of it may be collected into a globule at the bottom of the liquid. But in the meantime a portion of the copper will have been

dissolved, and has now taken the place formerly occupied by the mercury, and the solution is found to be green. The copper has, in fact, *displaced*, or been *substituted for* the mercury in the original compound, so that we have now in solution copper chloride instead of mercuric chloride. For every 100 grains of mercury which are separated in the metallic form, 31 grains of copper are dissolved, so that we learn that 31 parts of copper will displace 100 of mercury.

When the whole of the mercury has been separated, the clear green solution may be poured into another glass and the bright blade of a knife immersed in it. The iron will displace the copper, just as the copper displaced the mercury, an iron chloride (called ferrous chloride) remains in solution, and the copper is precipitated on the unaltered iron, giving it at first the appearance of being converted into copper. The 31 parts of copper which displaced 100 of mercury are themselves displaced by 28 of iron. Examples of this kind are very common in chemistry.

QUANTITIES CONCERNED IN CHEMICAL CHANGES.

It will readily be understood from the foregoing illustrations that it is of the utmost importance in studying chemistry to attend to the *relative quantities* of different substances which enter into chemical changes. It is to the careful study of these quantities that the dignity of chemistry as an exact science is due.

Law of constant composition.—This, the fundamental law of chemistry, may be stated in these words:

A given chemical compound always contains the same elements in the same proportions.

The law scarcely requires explanation. Water always contains, in 9 parts by weight, 8 of oxygen and 1 of hydrogen; mercuric oxide always contains 8 of oxygen to 100 of mercury; mercuric iodide, 100 of mercury to 127 of iodine, and so on. If 100 grains of mercury are heated with 130 grains of iodine, 227 grains only of mercuric iodide are produced, and 3 grains of iodine remain unchanged.

Difference between combination by weight, and by volume.—

Combination by volume means combination by measure, and it will be obvious to every one that, as bodies differ very much in weight, there will generally be a great difference between the proportions by measure and the proportions by weight in which two bodies combine. Take the case of the two gases hydrogen and chlorine (both elements). One volume or measure of hydrogen will combine with one volume of chlorine; but, as chlorine is $35\frac{1}{2}$ times heavier than hydrogen, one part by weight of hydrogen really combines with $35\frac{1}{2}$ parts by weight of chlorine.

In the cases of solids, and to a great extent of liquids also, we are compelled to trust almost entirely to the evidence of weight, for solids are very difficult to measure, and both solids and liquids exhibit such irregularities in regard to the volume that they occupy, that it has not been found possible to frame general laws for them. But with gases it is not so. They are very much easier to measure than to weigh; and as they exhibit a remarkable simplicity in the proportions by volume in which they act on one another; as moreover they are all equally affected in volume by alterations of temperature and pressure (Chap. II.), they are always measured for the purposes of experiment. Their weight (or specific gravity) is determined once for all with the utmost care, and it is then easy to calculate in a few minutes the weight of any given volume of any one of them.

CHEMICAL ACTION BETWEEN GASES.—CONSTITUTION OF GASES.

*When chemical action takes place between two gases, the volumes of those gases which are concerned in the change are found to bear a definite and generally a very simple proportion to one another, and also to the volume of the resulting gas, if the resulting compound be a gas.**

This fundamental law, which is sometimes called the

* It will be understood hereafter that the apparent exceptions to this rule are in perfect accordance with it. They only occur when a compound gas of complex structure is decomposed during a chemical change.

"law of gaseous volumes," will be better understood with the aid of a few illustrations. In studying them it is necessary to bear in mind the fact mentioned before, that all true vapours are gases. In considering the laws of gases, we have to include water-gas (steam), sulphur-gas, and even zinc-gas, for zinc is a gas at a high temperature, as well as oxygen and hydrogen, which are gases at ordinary temperatures. We deal in fact with matter in the gaseous state. For the sake of exact comparison all gases and vapours are assumed to be measured at a temperature of 32° F. (0° C.) and a pressure of 30 inches (760 millimetres) of mercury. Even if, as is the case with steam, the gas cannot really be cooled to 32° without condensing, it is easy to calculate by the formula already given (page 34) what its volume would be if it could.

1 volume of chlorine gas combines with 1 volume of hydrogen gas to form 2 volumes of hydrochloric acid gas. This is the simplest case.

2 volumes of hydrogen combine with 1 volume of oxygen to form 2 volumes of water-gas, or steam. Here 3 volumes condense, during the act of combination, into 2.

1 volume of arsenic vapour combines with 3 volumes of oxygen to form 1 volume of the vapour of white arsenic (arsenious anhydride). Here 4 volumes become condensed into 1.

These three cases will serve to illustrate the great simplicity that exists in the proportions in which gases unite by measure. Any given measure, pints, cubic inches, or litres, may of course be substituted for volumes in the above examples, and we may, for instance, say that every 2 pints of steam contain 2 pints of hydrogen and 1 pint of oxygen condensed into 2.

But this is not all. When a gas is decomposed into other gases, the volumes of its separated constituents are all found to bear simple and definite proportions to that of the original gas.

The gas called nitrous oxide, or laughing gas, is composed of nitrogen and oxygen, but it cannot be formed by the direct combination of these two elements. Nevertheless it is found, that when it is decomposed every 2 volumes yield 2 volumes of nitrogen and 1 volume of oxygen, so

that there can be no doubt that its constitution is similar to that of steam.

Ammonia-gas, again, is composed of nitrogen and hydrogen. When it is decomposed by electric sparks every 2 volumes yield 1 volume of nitrogen and 3 of hydrogen. The gases when separated from one another occupy in fact twice the volume that they did in a state of combination.

Standard of Volume for Gases.—The next point that we have to determine is, how best to represent the constitution of the numerous gases which are known to us; or in other words, how to represent the kind and the quantity of the elements which are present in equal volumes of them. It is obviously necessary, if we would compare them, to deal with equal volumes. It would appear at first sight that the best and simplest plan would be to state in all cases the composition of *one volume* of each gas. It is quite possible to do so, and indeed it does not very much matter what volume we take as the standard. But if we select one volume we are compelled to make constant use of fractions in representing the composition of gases, and it is therefore usual to take two *volumes* as the standard of comparison. A single example will make this plain.

We have already seen that 1 volume of chlorine combines with 1 volume of hydrogen to form 2 volumes of hydrochloric acid. If we wish to state the composition of 1 volume of hydrochloric acid, we must do so by saying that it contains $\frac{1}{2}$ a volume of chlorine and $\frac{1}{2}$ a volume of hydrogen. But on the 2-volume system the fractions are avoided. 2 volumes of chlorine are said to combine with 2 of hydrogen to form 4 of hydrochloric acid; and 2 volumes of hydrochloric acid contain 1 of chlorine and 1 of hydrogen.

Let us therefore agree to state in every case the composition of 2 volumes of gas. If we do so, we must never represent less than 2 volumes as taking part in any chemical process. We must, in fact, assign arbitrarily to the smallest quantity of any gas that is ever concerned in any chemical change a volume of 2. It is almost needless to add that our facts will then be equally true if we apply them to 2 millionths of a cubic inch or to 2 gallons.

Illustrations by Absolute Weights and Measures.—It will

occasionally be convenient in this study to supply illustrations in absolute weights and measures. We may do this in simple numbers in the following manner.

11.2 litres of hydrogen (11.16 more exactly) weigh 1 gramme. 11.2 litres may therefore be taken as 1 volume, and 22.4 litres as 2 volumes. As the specific gravity of a gas is the number of times it is heavier than hydrogen, *the specific gravity of every gas, taken in grammes, has a volume of 11.2 litres.* For example, as 11.2 litres of hydrogen weigh 1 gramme, 11.2 litres of oxygen weigh 16 grammes, since oxygen is 16 times as heavy as hydrogen. This is a very useful fact to remember, and it may be stated in another way. 1 litre of hydrogen weighs 0.09 gramme (0.0896 more exactly). Therefore, to find the weight of 1 litre of any gas, multiply its specific gravity by 0.09.

Chemical Symbols and Formulæ.—Each one of the elements is denoted in the system of modern chemistry by a *symbol*, which is either the first letter, or the first and some other characteristic letter, of its *Latin* name. Thus, oxygen is denoted by O; hydrogen by H; carbon by C; chlorine, which begins with the same letter, by Cl; calcium by Ca, and so on. In most cases the Latin name is the same as the English, but in a few instances it is different and the symbol must be remembered carefully. The following are the exceptions: Antimony is Sb, from stibium; copper, Cu, from cuprum; gold, Au, from aurum; iron, Fe, from ferrum; lead, Pb, from plumbum; mercury, Hg, from hydrargyrum; potassium, K, from kalium; silver, Ag, from argentum; sodium, Na, from natrium; tin, Sn, from stannum; and tungsten, W, from wolfram. Compounds are described by *formulæ*, which consist of the symbols of the elements composing them. Thus the formula for hydrochloric acid is HCl, for mercuric oxide, HgO, and so on.

Symbols then are employed to denote the elements, but they are also employed, by a very useful extension, to denote *definite weights of them*. What these weights are we have now to consider, confining ourselves for the present exclusively to elements in the state of gas.

To begin with, let us represent the standard 2 volumes

of hydrogen by the symbol H, and 2 volumes of chlorine by the symbol Cl. How are we in this case to represent hydrochloric acid, which, as we have already seen, contains in 2 volumes, 1 of hydrogen and 1 of chlorine? 2 volumes of this compound must evidently be represented by such a formula as this: $H\frac{1}{2} Cl\frac{1}{2}$. This is inconvenient, and yet such cases will be of incessant occurrence, unless we alter our representation of the elements. Accordingly we describe the 2 volumes of hydrogen as H_2 and the 2 volumes of chlorine as Cl_2 , and the formula for hydrochloric acid, which contains $\frac{1}{2}$ of each, then becomes HCl.

To take another instance. The symbol N might be applied to 2 volumes of nitrogen, if we did not know that 2 volumes of ammonia contain only 1 volume of nitrogen united to 3 volumes of hydrogen. If we use the symbol N for 2 volumes of nitrogen and H for 2 volumes of hydrogen, ammonia must be described by the inconvenient formula $N\frac{1}{2} H1\frac{1}{2}$. But calling nitrogen N_2 , and hydrogen H_2 , both of these are doubled, and the formula for ammonia becomes NH_3 .

Again, with phosphorus: we might apply the symbol P to 2 volumes of phosphorus-vapour if it were not for such compounds as phosphine gas, 2 volumes of which contain only one quarter as much phosphorus as is present in 2 volumes of phosphorus-vapour. But by describing phosphorus as P, and phosphine as PH_3 , we get over the difficulty; for the formulæ show plainly, and in whole numbers, that one contains four times as much phosphorus in 2 volumes as the other.

The symbol therefore denotes the least weight of the element formed in 2 volumes of gas.

Relative Weight of Gases.—It has already been shown that the specific gravity of a body means its weight as compared with the weight of an equal volume of some other body which is taken as a standard. The standard for gases is hydrogen, 1 volume of which is said to weigh 1. We have therefore only to double the specific gravities of the gases (which have been carefully determined by experiment) to find the relative weights of the standard 2 volumes of each gas. If 2 volumes of hydrogen weigh 2 (2 grains, 2 grammes, or 2 hundred-

weight), 2 volumes of chlorine will weigh 71, 2 volumes of oxygen, 32, and so on with the rest. We must now study the relative weights of the different elements which go to make up the weight of 2 volumes of each of the more important gases, elementary and compound. It will be best to take the elementary gases first.

Constitution of Elementary Gases.—It has already been seen that the formula for 2 volumes of any elementary gas is determined by the compounds which contain that gas. The formula for hydrogen is H_2 , because compounds are known which contain in 2 volumes only half as much hydrogen as is contained in 2 volumes of the pure element. Now what is true of a large volume must also be true of any volume, however minute, and we are therefore led to the conclusion that even the smallest conceivable volume of hydrogen gas that exists in a separate state must be capable of division into two parts, and as the weight of 2 volumes of hydrogen has already been defined to be 2 (hydrogen being the standard), each of these constituent parts must have a weight of 1.

*Atoms.**—The name *atom* is, for reasons that will be explained hereafter, applied to the quantity of each element that is denoted by its symbol. 2 volumes of hydrogen denoted by the formula H_2 , weigh 2, and consist of 2 *atoms*, each denoted by the *symbol* H, and each weighing 1. In like manner, 2 volumes of phosphorus-vapour are said to consist of 4 atoms; the formula for 2 volumes being P_4 and the symbol for each atom P. The atom in this case weighs 31. In the case of mercury and a few other elements the 2 volumes of gas are said to contain but 1 atom (weighing, in the case of mercury, 200), because no gas is known which contains in 2 volumes less than 200 of mercury.

The atom is therefore the smallest weight of each

* I use this word in its present connection with regret. It would certainly be better to exclude it altogether from an account of gases which is independent of the truth or untruth of the atomic theory. But there is no word in use which exactly answers to the conditions, and I hold the coinage of a word to be far too serious an experiment to be undertaken in an elementary treatise. Might not the word "prime," originally employed by Wollaston, be conveniently revived for this purpose?

element ever found in 2 volumes, of any gas, the smallest weight of hydrogen so found being taken as 1. The weights of the atoms are, of course, relative weights. By the use of small index figures any number of atoms may be represented. Thus S_6 means 6 atoms of sulphur each weighing 32.

These remarks will sufficiently explain the following table which exhibits the composition of all the more important elementary gases. It is only necessary to point out that three of them, oxygen, sulphur, and iodine, occur in two places. Two different modifications of each of these gases are known, which differ from one another in weight. The *atoms* are, however, the same in each case.

To translate the numbers in the three following tables into absolute weights and measures, the system described on page may be adopted. For 2 volumes, read 22.4 litres. Then all the weights will be grammes. Thus in the first table 22.4 litres of hydrogen weigh 2 grammes, of oxygen, 32 grammes, of ozone, 48 grammes, and so on.

Constitution of Compound Gases.—It will now be obvious that two volumes of every compound gas must contain two or more atoms of two or more kinds. The formula for a compound gas must denote the number and kind of atoms that there are in two volumes of it. Hydrochloric acid is an example of the simplest kind of gaseous compound, two volumes of it containing, as we have already seen, one atom of hydrogen and one of chlorine. Other gases contain atoms of three and even four different elements, and there are sometimes as many as twenty, thirty, or even more atoms in two volumes. In fact, we know no limit to the number of atoms which may enter into the composition of a compound gas.

The following table shows the constitution and formulæ of a few of the more important compound gases. The rules according to which they are named will be explained hereafter.

Compound Gases containing non-volatile Elements.—Many gases contain atoms of elements which have not as yet been converted into the condition of gas, or which, even if they do exist as gas, have never had their specific

ELEMENTARY GASES.

NAME OF GAS.	Sp. Gr. or weight of 1 volume.	Weight of 2 volumes.	No. of atoms in 2 volumes.	Symbol for each atom.	Weight of each atom.	Formula for 2 volumes.
Mercury	100	200	1	Hg	209	Hg
Zinc	32.5	65	1	Zn	65	Zn
Cadmium	56	112	1	Cd	112	Cd
Iodine	63.5	127	1	I	127	I
Hydrogen	1	2	2	H	1	H ₂
Chlorine	35.5	71	2	Cl	35.5	Cl ₂
Bromine	80	160	2	Br	80	Br ₂
Iodine	127	254	2	I	127	I ₂
Oxygen	16	32	2	O	16	O ₂
Sulphur (above 1000° C.)	32	64	2	S	32	S ₂
Nitrogen	14	28	2	N	14	N ₂
Oxygen (as ozone) . .	24	48	3	O	16	O ₃
Phosphorus	62	124	4	P	31	P ₄
Arsenic	150	300	4	As	75	As ₄
Sulphur (below 1000° C.)	96	192	6	S	32	S ₆

gravities accurately determined in that condition. In these cases the atomic weight (weight of the atom) of the non-volatile element can be found by exactly the same rule as that given before, namely, by observing the smallest weight of the element that ever enters into the composition of 2 volumes of a compound gas.

Take the case of carbon, an element which has never been converted into vapour. Many gases containing carbon are known, but not one of them contains in 2 volumes less than 12 of carbon. The atomic weight of carbon is therefore taken to be 12 and the atom is denoted by the symbol C, but not knowing the weight of 2 volumes of carbon-vapour we cannot tell what fraction 12 is of that weight; or, in other words, how many atoms of carbon there are in 2 volumes of carbon-gas.

It is obvious that we cannot have a *formula* for a non-volatile element, and that even in the case of volatile elements it is incorrect to apply the formula belonging to

the gas to the liquid or solid element. For liquid and solid elements we must be content to use the *symbol* which denotes the atom.

COMPOUND GASES.

NAME OF GAS.	Specific gravity.	Weight of 2 volumes	Kind, number, and weight of the atoms in 2 volumes.	Formula.
Mercuric Chloride .	135.5	271	{ Mercury 1 = 200 Chlorine 2 = 71 }	Hg Cl ₂
Hydrochloric Acid .	18.25	36.5	{ Hydrogen 1 = 1 Chlorine 1 = 35.5 }	H Cl
Hypochlorous An- hydride . . . }	43.5	87	{ Oxygen 1 = 16 Chlorine 2 = 71 }	Cl ₂ O
Water. (Steam) .	9	18	{ Hydrogen 2 = 2 Oxygen 1 = 16 }	H ₂ O
Hydrogen Sulphide.	17	34	{ Hydrogen 2 = 2 Sulphur 1 = 32 }	H ₂ S
Sulphurous Anhy- dride . . . }	32	64	{ Oxygen 2 = 32 Sulphur 1 = 32 }	S O ₂
Ammonia . . .	8.5	17	{ Hydrogen 3 = 3 Nitrogen 1 = 14 }	H ₃ N
Nitrous Oxide . .	22	44	{ Nitrogen 2 = 28 Oxygen 1 = 16 }	N ₂ O
Nitric Oxide . .	15	30	{ Nitrogen 1 = 14 Oxygen 1 = 16 }	N O
Phosphine . . .	17	34	{ Hydrogen 3 = 3 Phosphorus 1 = 31 }	H ₃ P
Phosphorous Chloride	68.75	137.5	{ Chlorine 3 = 106.5 Phosphorus 1 = 31 }	Cl ₃ P
Arsine	39	78	{ Hydrogen 3 = 3 Arsenic 1 = 75 }	H ₃ As

The following table will show the constitution of a few important gases of the kind here specified and also the way in which the atomic weights of their non-volatile constituents can be determined from them.

The atomic weight of iron, as deduced from the composition of the vapour of ferric chloride, is 112. But

ferric chloride is so similar to chromic chloride that the atomic weight of iron is held to be 56. The formula for ferric chloride is then Fe_2Cl_6 ; 2 atoms of iron, each 56, united with 6 atoms of chlorine, each 35.5. This brings it into analogy with chromic chloride Cr_2Cl_6 .

COMPOUND GASES CONTAINING NON-VOLATILE ELEMENTS.

NAME OF GAS.	Specific gravity.	Weight of 2 volumes	Kind, weight, and number of the atoms in 2 volumes.	Formula.	Symbol and atomic weight of the non-volatile constituent.
Methane .	8	16	{ Hydrogen 4 = 4 Carbon 1 = 12 }	CH_4	C = 12
Ethylene .	14	28	{ Hydrogen 4 = 4 Carbon 2 = 24 }	C_2H_4	
Ethane .	15	30	{ Hydrogen 6 = 6 Carbon 2 = 24 }	C_2H_6	
Carbonic oxide }	14	28	{ Oxygen 1 = 16 Carbon 1 = 12 }	CO	
Carbonic anhydride }	22	44	{ Oxygen 2 = 32 Carbon 1 = 12 }	CO_2	
Cyanogen .	26	52	{ Nitrogen 2 = 28 Carbon 2 = 24 }	C_2N_2	
Chloroform	59.75	119.5	{ Hydrogen 1 = 1 Chlorine 3 = 106.5 Carbon 1 = 12 }	CHCl_3	
Silicon chloride }	85	170	{ Chlorine 4 = 142 Silicon 1 = 28 }	SiCl_4	Si = 28
Methyl silicate }	76	152	{ Carbon 4 = 48 Hydrogen 12 = 12 Silicon 1 = 28 Oxygen 4 = 64 }	$\text{C}_4\text{H}_{12}\text{SiO}_4$	
Chromyl chloride }	77.75	155.5	{ Oxygen 2 = 32 Chlorine 2 = 71 Chromium 1 = 52.5 }	CrO_2Cl_2	
Chromic chloride }	159	318	{ Chlorine 6 = 213 Chromium 2 = 105 }	Cr_2Cl_6	Cr = 52.5
Ferric chloride }	162.5	325	{ Chlorine 6 = 213 Iron 1 = 112 }	FeCl_6	Fe = 112

The composition and properties of other iron compounds confirm this view.

Calculation of the Specific Gravities of Gases from their Formulæ.—If the formula for any gas be accurate; that is, if it truly represents 2 volumes of the gas, the specific gravity of the gas may be calculated from the formula. This may be illustrated by an example.

The formula for carbonic anhydride gas is CO_2 . What is its specific gravity?

The formula tells us that 2 volumes of the gas contain 1 atom of carbon, the weight of which we know to be 12, and 2 atoms of oxygen, each of which weighs 16, total 44. 2 volumes of carbonic anhydride therefore weigh 44, whereas 2 volumes of hydrogen weigh 2; which shows that carbonic anhydride is 22 times heavier than hydrogen; or, in other words, that its specific gravity is 22.

To find the specific gravity of any gas, elementary or compound, we have therefore only to find the weight of 2 volumes of it, by adding together the weights of the constituent atoms, and then to divide the number so obtained by 2.

COMPOSITION AS DETERMINED BY WEIGHT.—LAWS OF COMBINATION BY WEIGHT.

We must now leave on one side for the present all considerations of volume and confine ourselves to the examination of the *weights* of different substances which are concerned in chemical changes, and the constitution of compounds as determined by weight. This method of study is of universal application, and it is used indifferently for solids, liquids, and gases, whereas we have already seen that the study of volumes only gives satisfactory results when applied to gases.

Percentage Composition.—The most obvious way of stating the composition of a compound is as parts in 100. The results of an analysis are always calculated in this way first of all, the figures being usually carried to the second or third decimal place. Thus, if 20 grammes of lime are analysed they are found to contain 14.286 grammes of

calcium, 5.714 of oxygen. By a simple proportion sum it is then found that the percentage composition is :

Calcium	71.43
Oxygen	28.57
	100.00

For 20 : 100 :: 14.286 : 71.43
 And 20 : 100 :: 5.714 : 28.57

The percentage composition of a few important hydrogen compounds is shown in the following table.

Simplest numerical Proportion of the Constituents.—From the percentage composition, it is easy to calculate the pro-

NAME OF COMPOUND.	Composition of 100 parts.	Composition. H=1.
Hydrochloric acid .	{ Hydrogen 2.74 Chlorine 97.26	1 35.5
Water	{ Hydrogen 11.1 Oxygen 88.8	1 8
Ammonia	{ Hydrogen 17.65 Nitrogen 82.35	1 4.6
Methane	{ Hydrogen 25. Carbon 75.	1 3
Ethylene	{ Hydrogen 14.29 Carbon 85.71	1 6
Acetylene	{ Hydrogen 7.69 Carbon 92.31	1 12
Hydrogen sulphide .	{ Hydrogen 5.88 Sulphur 94.12	1 16

portion that the weight of each constituent bears to that of some one which is taken as unity. In the preceding table the last column shows the weights of various elements which are combined with 1 of hydrogen in a few important compounds, the proportion being of course the same as in the percentage composition. Any other element might be taken instead of hydrogen as the standard. It

was indeed common at one time to take oxygen as the standard, calling it 100, but hydrogen is more convenient and is now generally employed.

When the composition of hydrogen compounds is represented in this way, very simple numbers are for the most part obtained, and some numerical relations are observed which are hidden in the mere percentage compositions. Thus in the three carbon compounds (which are examples of a large number actually known) we observe that 1 of hydrogen unites with 3, 6, and 12 of carbon. 3 is the smallest proportion ever found.

By extending the above table we can easily obtain a series of numbers which represent the proportions in which a good many of the most important elements combine with one part of hydrogen. It is now necessary to extend our study to those compounds which do not contain hydrogen.

And here there are two courses open to us. We may begin by choosing some other element as a standard, calling its quantity 1, and calculating the weight of other elements which combine with it. But this method, though it serves to exhibit some interesting relations, does not bring out the simple general law towards which we are tending. We must make the comparison in another way.

Knowing the weight of any element which unites with 1 of hydrogen, let us calculate the weight of other elements which unite *with that weight*. For example: 35.5 of chlorine unite with 1 of hydrogen. Instead therefore of calculating how much of each other element will unite with 1 of chlorine, let us calculate how much will combine with 35.5 of chlorine. And inasmuch as 8 of oxygen combine with 1 of hydrogen, let us see what weight of each other element unites with 8 of oxygen. In examining the results so obtained, it must be remembered that two elements often unite in several different proportions.

Confining ourselves to the elements mentioned on the preceding table, we can obtain, from the analyses of well-known compounds, the following figures, which are merely samples of an immense number at our disposal:

35.5 parts by weight of chlorine unite with 1 of hydrogen; with 8, 24, and 32, of oxygen; with 3, 4, 6, and 12, of carbon, and with 32 of sulphur.

8 parts of oxygen unite with 1 and with 0.5 parts of hydrogen; with 8.875, 11.83, and 35.5 parts of chlorine; with 2.8, 3.5, 4.6, 7 and 14 parts of nitrogen; with 3 and 6 parts of carbon, and with 5.3 and 8 parts of sulphur.

4.6 parts of nitrogen unite with 1 of hydrogen; with 2.6, 5.3, 8, 10.6, and 13.3 parts of oxygen, and with 4 parts of carbon.

3 parts of carbon (the smallest quantity that unites with 1 of hydrogen) unite with 0.25, 0.5 and 1 part of hydrogen; with 8.875, 17.75, 26.625, and 35.5 of chlorine; with 4 and 8 parts of oxygen; with 7 of nitrogen, and 16 of sulphur.

16 parts of sulphur unite with 16 and 24 of oxygen; with 17.75 of chlorine and with 3 of carbon.

A careful study of these figures will elucidate two of the most important laws of chemistry. It will be seen:—

1. *That in some cases the proportions in which two elements combine with one of hydrogen are exactly the proportions in which they unite with one another.* 35.5 parts of chlorine and 8 parts of oxygen will respectively combine with 1 of hydrogen, and a compound is known which contains in every 43.5 parts, 35.5 of chlorine and 8 of oxygen.

2. *That even when two elements unite in proportions different from those in which they unite with one of hydrogen, the numbers representing the two proportions bear a simple relation to one another.*—Chlorine forms three compounds with oxygen. The one mentioned above contains 35.5 of chlorine to 8 of oxygen; the other two, 35.5 chlorine to 24 and 32 of oxygen. In these two latter the proportion of oxygen is exactly three and four times as great as is found combined with hydrogen in water. The five oxides of nitrogen afford a still more striking example, although some thought is required to make it apparent. We already know that 1 of hydrogen unites with 4.6 of nitrogen. In the following table the columns next to the names of the compounds show the quantities of oxygen which are combined with 4.6 of nitrogen in each of the oxides. These quantities of oxygen are represented by somewhat

awkward fractions. But if we multiply the 4.6 by 3, and note the quantities of oxygen which unite with 14 of nitrogen, as in the middle columns of the table, we find that the proportions of oxygen can be expressed by whole numbers, and furthermore that these numbers are simple multiples of 8, which we already know is the proportion in which oxygen unites with 1 of hydrogen. For the sake of comparison the percentage compositions are given in the last columns.

OXIDES OF NITROGEN.

					Parts in 100.	
	Nitrogen.	Oxygen.	Nitrogen.	Oxygen.	Nitrogen.	Oxygen.
Nitrous oxide	4.6	2.6	14	8	63.6	36.4
Nitric oxide	4.6	5.3	14	16	46.7	53.3
Nitrous anhydride	4.6	8.0	14	24	36.8	63.2
Nitric peroxide	4.6	10.6	14	32	30.4	69.6
Nitric anhydride	4.6	13.3	14	40	25.9	74.1

Units of Combining Weight.—*Combining Weights.*—*Atomic Weights.*—By an extension of the above methods of comparison, it is at last made out that a number may be found for each element which is variously known as its “combining weight,” “proportional number,” or, for reasons which will be explained hereafter, its “atomic weight.” This number denotes the smallest quantity which is ever found united with one of hydrogen, or with an analogous quantity of any other element. For shortness it may be spoken of as the *atom* of the substance. When we have once fixed upon the numbers which shall stand for the atoms of the elements, it is easy to represent the composition of all compounds by stating the number and kind of atoms which they contain. Thus with the oxides of nitrogen. If we agree that the atom of nitrogen is 14 and that of oxygen 8, we see by the above table that in the 5 oxides, one atom of nitrogen is united with 1, 2, 3, 4 and 5 atoms of oxygen.

Use of Symbols.—The symbols which have already been used to denote volumes of gas, may be applied with equal exactness of meaning to the atoms of elements as found by weight. Thus, H may denote 1 atom, or 1 part by weight of hydrogen. N, 1 atom, or 14 parts of nitrogen. O, 1 atom, or 8 parts of oxygen, and so on. With these symbols, formulæ, perfectly analogous to those already

explained, may be constructed and applied to compounds. The formulæ for the 5 oxides of nitrogen are, if the above mentioned atomic weights be adopted: NO , NO_2 , NO_3 , NO_4 , and NO_5 .

Modes of fixing the Atomic Weights of Elements.—The numbers found for the combining or atomic weights of the elements from the study of the percentage composition of their compounds are liable to one serious drawback. We can never be certain that the number adopted represents the weight of any real unit of the element. We cannot indeed by fair reasoning conclude that there is any constant unit of combining weight for each element. All that we can say is, that if there be such a unit, it must either have the weight we have assigned to it or bear some simple numerical relation to that weight. Looking at the carbon compounds, for example, in the table on page 76, we come to the conclusion that the atomic weight of carbon is either 3, 6, or 12, for carbon unites with 1 of hydrogen in all those proportions. The atomic weight of nitrogen indicated by its hydrogen compound is 4.6, but the oxygen compounds of the same element seem to require it to be 14. In the same way the atomic weight of oxygen may be 4, 8, 16, or even some less simple number, and in all these cases analysis is powerless to tell us with any certainty which of the different numbers should be adopted. In fact, to speak accurately, analysis would appear to indicate that each element has more than one, or indeed many combining weights, but that these weights all bear a simple relation to one another.

Fortunately, however, there are other facts at our disposal with regard to the elements of their compounds which give clear evidence upon this point, and which not only indicate that there is a single unit of combination for each element, but enable us to fix its weight with a near approach to certainty. The chief of these exterior sources of evidence may be briefly mentioned in this place.

1. *The Law of Gaseous Volumes*, with the extensions of it that have already been considered. This method can of course only be applied where gases are concerned, but its evidence when it can be obtained is exceedingly valuable. The weights of the atoms of a great many of

the elements are fixed in this way, and the method has been so fully explained that it is unnecessary to do more than refer to it here. We have seen, to take a single example, that the atomic weight of oxygen is taken as 16, because 16 is the smallest weight of oxygen that occurs in 2 volumes of any gas.

2. *Substitution.—Equivalents.*—This is merely a modification of the method by weight which has already been described. It has been seen (page 63) that one of the commonest modes in which chemical action takes place is by the *substitution* of a certain weight of one substance for another. In a series of substitutions we have only to ascertain the weight of different substances, which will displace one another to obtain a series of numbers which are called the *equivalents* of those substances. Starting from hydrochloric acid, for instance, which contains 1 of hydrogen to 35.5 of chlorine, we find we can replace the 1 of hydrogen by 23 of sodium, 39 of potassium, 20 of calcium, and so on. And in the same way the 35.5 of chlorine may be replaced (either directly or indirectly) by 8 of oxygen, 4.6 of nitrogen, 80 of bromine, 16 of sulphur, and so on. These various numbers are the *equivalents* of the elements, and it is found that the quantities they represent are not only equivalent to 1 of hydrogen, or 35.5 of chlorine, but also to one another. 16 of sulphur may take the place of 8 of oxygen or 80 of bromine, and 20 of calcium are always equivalent to 39 of potassium.

All this is simple enough, and it would seem as if the equivalents so found might be adopted as the most convenient of "combining weights." They were in fact adopted, and for years were pretty generally employed for this purpose. But certain difficulties and inconveniences attend their use and have at last occasioned their abandonment. For, in the first place, the equivalents of some elements do not agree with the numbers indicated by the law of volumes. The equivalent of oxygen, for instance, is 8, but its atom, as deduced from the composition of compound gases, is 16. A still graver objection lies in the fact that as elements often combine in more than one proportion, it then follows that there

are several equivalents for the same element. The case of the oxides of nitrogen, already cited, will make this plain. On page 76, it is stated that 8 parts of oxygen will combine with 1 of hydrogen, and with 2.8, 3.5, 4.6, 7 and 14, of nitrogen; it is therefore evident that these five quantities of nitrogen are all, in different compounds, equivalent to 1 of hydrogen, and that, in fact, nitrogen has five equivalents. This difficulty is indeed only another form of that which was pointed out at the beginning of this section.

But although the equivalent values of the elements are no longer thought sufficient to fix their atomic weights, this method of study is still of immense importance in chemistry, for two reasons:

First, because, as we have seen, the true atomic weight, when not identical with the equivalent, always bears a simple relation to it.

And secondly, because it often enables to fix the exact constitution of a compound, as will be seen from the following simple illustration. The formula deduced from the analysis of acetic acid is CH_2O ($\text{C} = 12$, $\text{H} = 1$, $\text{O} = 16$). But we find by experiment that one quarter of the hydrogen of acetic acid may be replaced by a metal—by sodium, for instance, or silver. Now it is evident that, if the above formula were correct, the metal must displace *half an atom* of hydrogen, which, as the atom is defined to be the smallest possible quantity, is absurd. We must, therefore, double the formula and make it $\text{C}_2\text{H}_4\text{O}_2$. It is then seen to consist of 2 atoms of carbon, 4 of hydrogen, and 2 of oxygen. Sodium displaces one atom of hydrogen, and the formula for the compound so obtained is $\text{C}_2\text{H}_3\text{NaO}_2$. This view is completely confirmed by the law of volumes. By the rule given on page 75, the specific gravity of the vapour of acetic acid ought, if the formula CH_2O were correct, to be 15, whereas experiment shows it to be 30.

3. *Specific Heat.—Atomic Heat.*—This method gives very good results with elements in the solid state, and is of the utmost importance. Equal weights of different substances heated to an equal extent, require very different times for cooling down to the same point. If one pound of water and one pound of mercury, at a temperature of 100°C .

(boiling point), be allowed to cool down to 50° , it will be found that the water is 30 times as long in doing so as the mercury, and gives out 30 times as much heat. Water at any one temperature contains 30 times as much heat as an equal weight of mercury at the same temperature, and analogous differences are observed with all other substances. The quantities of heat which equal weights of different bodies contain at the same temperature (measured by the time required for their cooling, or in some other way) are called their *specific heats*. The specific heat of water is greater than that of any other known substance (except hydrogen), and it is therefore taken as the standard, and called 1. The specific heat of liquid mercury is therefore 0.033.

Now if we compare together the specific heat of different elements in the solid state, we find, as the following table will show, that no relation can be observed between them. But if we compare together the specific heats, not of equal weights, but of *atomic weights*, which we can easily do by multiplying the specific heats of the first column by the atomic weights, we find that the numbers are very similar to one another, and approximate to 6.3. It is therefore inferred that if we could compare all elements in the solid state and at equal temperatures, in quantities proportionate to their atomic weights, and could avoid all errors of experiment, *they would all contain the same quantity of heat*. This quantity, found by multiplying the specific heat by the atomic weight, is called the *atomic heat* of the element.

It is easy to understand the assistance which this theory, discovered by Dulong and Petit (we must remember that it is only a theory, for its truth cannot be demonstrated), affords in fixing the atomic weights of elements. Let us take the case of zinc to illustrate it. The equivalent of zinc, as compared with hydrogen, is 32.5, and that number was formerly taken as the atomic weight of the metal. But the specific heat of zinc is 0.0955, and this number multiplied by 32.5 only gives 3.1, so that the atomic heat of zinc must be regarded as exceptional, and only half of the usual quantity. But if we take 65 as the atomic weight of zinc, the anomaly vanishes, and the atomic heat becomes 6.2. The atomic weights of many other

elements have of late years been doubled for the same reason.

Some few exceptions to this rule of atomic heat are known, of which carbon, boron and silicon are the most important, but at high temperatures these disappear.

It will be seen that if we wish to find the atomic weight of an element from its specific heat, we must divide the average atomic heat 6.3 by the specific heat. The atomic weight so obtained will not be exact, but it will serve to indicate whether a particular number or a fraction or multiple of it is to be chosen as the true atomic weight.

E.g. Atomic weight of tin is $\frac{6.3}{.0562} = 112$, which suffi-

SPECIFIC AND ATOMIC HEATS.

ELEMENT.	Specific Heat.	Atomic Weight.	Atomic Heat.
Bromine (solid) . .	0.0843	80.	6.7440
Iodine	0.0541	127.	6.8707
Potassium	0.1696	39.	6.6144
Sodium	0.2934	23.	6.7482
Silver	0.0570	108.	6.1560
Gold	0.0324	196.	6.3504
Phosphorus	0.1887	31.	5.8497
Arsenic	0.0814	75.	6.1050
Antimony	0.0508	120.	6.0960
Bismuth	0.0308	208.	6.3064
Sulphur	0.1776	32.	5.6832
Magnesium	0.2499	24.	5.9976
Zinc	0.0955	65.	6.2075
Mercury (solid) . .	0.0319	200.	6.3800
Copper	0.0951	63.5	6.0389
Lead	0.0314	207.	6.4998
Iron	0.1138	56.	6.3728
Aluminium	0.2143	27.	5.7861
Tin	0.0562	118.	6.6316
Platinum	0.0324	197.	6.3828

ciently indicates that 118, and not 59, should be taken as the atomic weight of tin.

4. *Isomorphism*.—It was discovered by Mitscherlich,

that compounds which might be supposed to have a similar chemical structure, very commonly crystallized in the same form, even though quite different in properties. Similar compounds of calcium, strontium, barium, and lead; similar sulphates, selenates, and chromates, and similar phosphates and arsenates, are examples of this law, and many others might be quoted. Compounds which exhibit this peculiarity are said to be *isomorphous* with one another (from *ἴσος*, like, and *μορφή*, form). Mitscherlich's law, though not universal, is yet so frequently true, that when two analogous substances yield crystals having the same angular measurement, there is a great *a priori* probability that their constituents are arranged in a similar manner, and this probability is sometimes of great use in fixing the atomic weight of an element. For example, the metal aluminium has only one oxide, alumina, which contains 18 of the metal to 16 of oxygen. As 16 is the atomic weight of oxygen, it would be natural to suppose that that of aluminium was 18, in which case the formula of alumina would be AlO . But alumina is isomorphous with red oxide of iron (ferric oxide), which is known to have the formula Fe_2O_3 , and it is therefore concluded that the formula of alumina must be Al_2O_3 , in which case the atomic weight of aluminium must be 27. Of course either formula agrees equally well with the result deduced from analysis. For if:—

$Al = 18$, then $AlO = 18$ Aluminium : 16 Oxygen.

$Al = 27$, then $Al_2O_3 = 54$ „ : 48 „

the proportion being the same in both cases. The number 27 is confirmed by the specific heat of aluminium, for—

$$\frac{6.3}{.2143} = 29.4$$

which is tolerably near.

The atomic weights of the principal elements, determined by the collation of all these various facts and theories, are given in the table at the commencement of this volume, to which is prefixed a definition of the terms.

EQUIVALENT VALUE OF ATOMS.—VALENCY, ATOMICITY, OR QUANTIVALENCE.

The above remarks will have made it evident that the quantity which is known as the atom of an element is not always the same as its equivalent. The equivalent of oxygen is 8, because 8 parts of oxygen will replace 35.5 of chlorine, or 1 of hydrogen, but the atomic weight of oxygen is held to be 16: the equivalent of nitrogen is 4.6, but its atomic weight is 14: and lastly, in its simplest compound, the equivalent of carbon is 3, whereas its atomic weight is 12. It follows from this, that an atom of oxygen is equivalent to two, an atom of nitrogen to three, and an atom of carbon to four atoms of hydrogen. For if 8 of oxygen be equivalent to 1 of hydrogen, 16 must evidently be equivalent to 2. The atom of each element has therefore its own equivalent value, represented by the number of atoms of hydrogen to which it is equivalent. This equivalent value, as compared with that of the atom of hydrogen, is often spoken of as the *valency*, or *atomicity*, or *quantivalence* of the element, and its amount is described by the words *monad*, *diad*, *triad*, *tetrad*, *pentad*, &c., according as the atom is equivalent to one, two, three, four, or five atoms of hydrogen.

Classification of Elements by their Valency.—All those elements of which one atom is equivalent to one atom of hydrogen, either in replacing it or combining with it, are classed together as *monads*. The most important monad elements are, hydrogen, chlorine, bromine, iodine, and fluorine among the non-metals, and potassium, sodium, and silver, among the metals.

The chief *diad* elements are oxygen and sulphur among the non-metals, and barium, strontium, calcium, magnesium, zinc, copper, mercury, and lead among the metals.

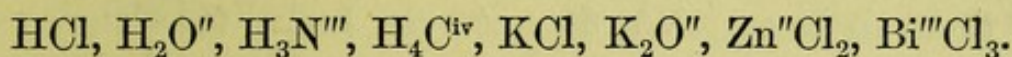
As examples of *triads* may be mentioned nitrogen, phosphorus, arsenic, and boron, and the metals antimony, bismuth and gold.

Carbon, silicon, and many of the other metals are *tetrads*.

Mode of marking Valency.—As the symbol denotes one atom, some chemists mark the equivalent value, or valency, of that atom by dashes or Roman figures placed above

and to the right of the symbol, in the following manner:— $H' O'' N''' C''''$ or C^{iv} . The monad atoms, however, are seldom marked at all.

Let it be understood distinctly that a diad atom may take the place of or may combine with any two monad atoms. One atom, or 16 parts of oxygen, for instance, may very often, by mere substitution, be made to *displace* and take the place of two atoms (2 parts) of hydrogen, and we already know that 16 parts of oxygen will *combine with* 2 of hydrogen. The formulæ for a few important compounds will illustrate the different values of the atoms more clearly than words.



Mode of fixing Valency.—When an element combines with hydrogen, the hydrogen compound decides the valency, as with chlorine, oxygen, nitrogen, and carbon. In other cases the chlorine compound, or any other compound that the element may form with a monad, may be used instead. Thus platinum is reckoned a tetrad because it forms the chloride, $Pt^{iv}Cl_4$, in which the atom of platinum is united with four of the monad atoms of chlorine.

Perissiad and Artiad Atoms.—Dr. Odling proposed to name those elements whose valency is an uneven number *perissiads* ($\pi\epsilon\rho\iota\sigma\sigma\acute{o}s$, odd), and those in which it is even, *artiads* ($\acute{\alpha}\rho\tau\iota\omicron s$, even). There is great convenience in the use of these words, but they are not often used.

Variations of Valency.—Many elements have different valencies in different compounds. For example, a second compound of platinum and chlorine is known, which has the formula $Pt''Cl'_2$, and here the metal is clearly a diad. So carbon, which is tetrad in $C^{iv}H_4$ and $C^{iv}O''_2$ (two diad atoms act, of course, like four monad ones), is diad in $C''O''$. And nitrogen, which is triad in $N'''H_3$, is monad in N'_2O'' , and pentad in N^vH_4Cl' . But these variations are subject to one rule, which is of very general application. *Perissiads are seldom artiads, or artiads perissiads.* A pentad, for example, which of course is a perissiad, may sometimes act as a triad, or a monad, but not as a tetrad, or diad. Two of the oxides of nitrogen, NO and NO_2 , are almost the only

certain exceptions of this important rule, which must be borne carefully in mind in constructing formulæ.

Radicals.—The formula for nitric acid is HNO_3 . Now a large number of compounds called *nitrates* are known, which differ from nitric acid only in containing some metal in place of hydrogen. Thus we have sodium nitrate NaNO_3 , and silver nitrate AgNO_3 . In all these nitrates, however, the common quantity NO_3 is found, and NO_3 is therefore called the *radical* of the nitrates. In the same way the formula for sulphuric acid is H_2SO_4 , and the sulphates all contain the radical SO_4 . Neither NO_3 nor SO_4 exist in a separate state. It is probably impossible for them to do so. We do not even know that they exist in the compounds, but we do know that their elements are present in the same proportions in a large number of compounds, and that all these compounds are linked together by a certain similarity of properties; and therefore, without troubling ourselves either to affirm or deny their existence as separate entities, it is convenient to assume their presence in certain compounds. We shall meet with many radicals as we go on, many of them very unlike one another in function; but for the present we have to regard them only in one point of view. In the compounds which contain them, *radicals play the part of elements*, and, like elements, every radical has its own proper valency. The acids which contain radicals afford good examples of this.

Nitric Acid	H N O_3	is like	Hydrochloric Acid	H Cl .
Sulphuric Acid	$\text{H}_2 \text{ S O}_4$	,,	Water	$\text{H}_2 \text{ O}$.
Phosphoric Acid	$\text{H}_3 \text{ P O}_4$	,,	Ammonia	$\text{H}_3 \text{ N}$.

As Cl is united with one of H, so NO_3 is united with one of H. NO_3 is therefore called a *monad radical*. And as SO_4 is united with two of H, it may be compared to oxygen, and is called a *diad radical*. Their valency may be marked as that of elements is marked $(\text{NO}_3)'$, $(\text{SO}_4)''$, and $(\text{PO}_4)'''$, and lastly it must be remembered that in all nitrates NO_3 is monad, and in all sulphates SO_4 is diad.

Measure of the Valency of Radicals.—The equivalent value of a radical is measured, just as that of an element is, by the number of atoms of hydrogen or other monad element with which it combines. NO_3 , for instance, is a

monad radical, because in HNO_3 it is combined with one atom of hydrogen. A radical is in fact an incomplete compound, and it is monad, diad, or triad according as it lacks one, two, or three monads to complete it. The following table includes some of the most important radicals of chemistry. Only a few of them have individual names.

MONAD RADICALS.

O H (Hydroxyl) in Hydrates	} as in H_2O . KOH . $\text{Ca}''(\text{OH})_2$. $\text{Bi}'''(\text{OH})_3$.
N H_2 (Amidogen) in Amides	} ,, NH_3 . KNH_2 .
CN (Cyanogen) in Cyanides	} ,, HCN . KCN . $\text{Hg}''(\text{CN})_2$.
Cl O in Hypochlorites	,, HClO . KClO .
Cl O_2 in Chlorites	,, HClO_2 . KClO_2 .
Cl O_3 in Chlorates	,, HClO_3 . KClO_3 .
Cl O_4 in Perchlorates	,, HClO_4 . KClO_4 .
N O_2 in Nitrites	,, HNO_2 . KNO_2 .
N O_2 in Nitro-compounds	} ,, $\text{C}_6\text{H}_5\text{NO}_2$.
N O_3 in Nitrates	,, HNO_3 . KNO_3 .
P O_3 in Metaphosphates	,, HPO_3 . KPO_3 .
N H_4 (Ammonium) in Ammonium Compounds, as in NH_4Cl . NH_4NO_3 .	
C H_3 (Methyl) in Methyl Compounds.	,, CH_3Cl .
C H_5 (Ethyl) in Ethyl Compounds	,, $\text{C}_2\text{H}_5\text{Cl}$.

DIADS.

S O_3 in Sulphites	as in H_2SO_3 .	K_2SO_3 .	$\text{Ca}''\text{SO}_3$.
S O_4 in Sulphates	,, H_2SO_4 .	K_2SO_4 .	KHSO_4 .
C O_3 in Carbonates	,, K_2CO_3 .	KHCO_3 .	$\text{Ca}''\text{CO}_3$.
Si O_3 in Metasilicates	,, H_2SiO_3 .	$\text{Mg}''\text{SiO}_3$.	

C H_2 (Methylene) in Methylene Compounds, as in CH_2Cl_2 .

TRIADS.

P O_4 in Phosphates, as in H_3PO_4 . K_3PO_4 . KH_2PO_4 . K_2HPO_4 .	
As O_4 in Arseniates	,, H_3AsO_4 . Ag_3AsO_4 .

C H in such compounds as Chloroform, CHCl_3 .

TETRADES.

Si O_4 in Silicates	as in H_4SiO_4 .	$\text{Mg}''_2\text{SiO}_4$.	
P O_7 in Pyrophosphates	,, $\text{H}_4\text{P}_2\text{O}_7$.	$\text{Na}_4\text{P}_2\text{O}_7$.	$\text{Mg}''_2\text{P}_2\text{O}_7$.

NOMENCLATURE.

Names are given to chemical compounds, according to rules which, though not perfect, are better than those

employed in any other science. Only a few of the simplest need be indicated in this place.

1. Compounds of two elements only are distinguished by the termination *ide*. Thus the compound of zinc and oxygen is called oxide of zinc, or zinc oxide; zinc and chlorine, zinc chloride; sodium and chlorine, sodium chloride; and so on. This rule is invariable, though for special reasons a different name is more often used for many simple compounds.

2. When there are two compounds of the same two elements, the lower—that which contains the least oxygen, chlorine, &c.—is distinguished by the termination *ous*, and the higher by the termination *ic*. The lower chloride of tin SnCl_2 , for instance, is called stannous chloride, and the higher, SnCl_4 , stannic chloride. The same rule is applied to acids. HNO_2 is called nitrous acid, and HNO_3 nitric acid.

3. The substances called salts which are derived from an acid which ends in *ous* have the termination *ite*. All the salts, for instance, which contain the radical NO_2 , and which therefore correspond to nitrous acid, are called nitrites. Salts which correspond to acids ending in *ic*, have the termination *ate*. Nitric acid, for instance, forms nitrates, sulphuric acid, sulphates, and acetic acid, acetates. Almost the only exception to this rule is in the compounds called hydrates or hydroxides, which contain the radical HO , and which, instead of corresponding in composition to an acid, correspond to water.

Other rules will be more conveniently given as instances occur.

ATOMIC AND MOLECULAR HYPOTHESES.

Atoms—Atomic Weights.—Dalton, to whom we are indebted for the first clear enunciation of the laws of combination by weight, contrived a beautiful hypothesis to account for them. In its simplest form it is known as Dalton's atomic hypothesis. According to this view, matter is composed of ultimate indivisible particles called *atoms* ($\alpha\tau\omicron\mu\omicron\varsigma$, indivisible). Every element has its own atom, peculiar in properties and in weight. All compounds are formed by the union together of elementary

atoms. The relative weights of the atoms are expressed by their combining weights, and afford a simple reason for those combining weights. Thus (taking the modern numbers), 35.5 parts of chlorine combine with 1 of hydrogen, because every atom of chlorine is attracted by and combines with one atom of hydrogen. In the same way, every one atom of oxygen, weighing 16, combines with two atoms of hydrogen, weighing 2. In fact, the combining weights of the elements show the relative weights of the atoms, hydrogen being taken as unity, and this is the reason why the combining weights are generally called atomic weights. It follows from this that the ultimate particle—the smallest possible quantity—of any compound must contain at least two different elementary atoms. Compounds are groups of atoms of two or more kinds.

Molecules.—Molecular Weights.—The ultimate particles of compounds, consisting, as we have seen, of groups of atoms, are called *molecules* (*moleculus*, small mass). Their weight is, of course, the sum of the weights of their constituent atoms. Thus the molecular weight of hydrochloric acid, HCl, is 36.5, and of water, 18. Reasons have already been given for believing that the ultimate particles of most elements, in the gaseous state, consist of two or more similar atoms. The view adopted for compounds must therefore be extended to elements, and we must admit that in the gaseous state elements consist of molecules incapable of division, which molecules themselves consist of atoms of similar kind, sometimes of only one (mercury, zinc, and cadmium gases), sometimes of two (hydrogen, oxygen, &c.), and sometimes of more than two (ozone, phosphorus, arsenic, and sulphur gases).

By an extension of this view, the symbols and formulæ which are used to denote two volumes of any gas, elementary or compound, can also be applied with equal accuracy to denote respectively *one atom* and *one molecule* of it. The atom of hydrogen may be denoted by H, the molecule by H_2 ; the atom of arsenic by As, the molecule (as gas) by As_4 ; and single molecules of hydrochloric acid, steam, and ammonia may be represented by the formulæ HCl, H_2O , and H_3N . This leads us to another

important hypothesis, which of late years has been very generally adopted. It refers only to gases.

Hypothesis of Avogadro and Ampere.—This hypothesis, suggested by Avogadro in 1811, and further developed by Ampere in 1814, may be stated in the following terms:

Equal volumes of different gases, at equal pressure and temperature, contain equal numbers of molecules.

This hypothesis explains two things.

1. The fact already mentioned (pages 34 and 57), that all gases are equally affected in volume by variations of pressure and temperature. Physical agencies, such as heat and pressure, do not affect the nature of molecules, but only their distance from one another, and this being the same in all gases under similar conditions, the effect produced by those agencies must naturally be the same in all cases.

2. The law of gaseous volumes. If one molecule of chlorine enters into chemical action with one molecule of hydrogen, the same will be true of one thousand, or one million, molecules of each. Now by the hypothesis one thousand, or one million, molecules of chlorine occupy a space equal to that occupied by one thousand, or one million, molecules of hydrogen, and it therefore follows that the two elements should react together in the proportion of equal volumes, which experiment proves to be the fact. In like manner, two volumes of hydrogen react with one volume of oxygen, because one volume of oxygen contains n , and two volumes of hydrogen $2n$ molecules, and every molecule of oxygen reacts with two molecules of hydrogen.

Structure of Gases.—Summary of hypotheses.—The modern doctrine may be summed up in these words.

1. *Equal volumes of gases, at equal pressure and temperature, contain equal numbers of molecules.*

2. *Every molecule consists of one or more atoms, either similar in kind (elementary molecules), or dissimilar in kind (compound molecules).**

* The molecular hypothesis is sometimes referred to so loosely, that a few further remarks upon it may not be out of place.

1. The hypothesis asserts nothing of the relative bulk of molecules, or atoms. It does not assert that one molecule of hydrogen has the

Molecules of Solids and Liquids.—The formulæ applied to gases are generally extended for convenience sake, but by imperfect reasoning, to the same substances in the liquid and solid state. But it is right to point out that neither hypothesis nor experiment tell us anything of the number of atoms contained in the molecules of solids and liquids. We have good reason for believing that a molecule of phosphorus-vapour contains four atoms; but we do not know how many a molecule of solid phosphorus may contain. The formula for common salt, which is non-volatile, is written NaCl , on account of its analogy, not with solid, but with gaseous HCl .

Hypothesis to account for Valency.—The different equivalent value of different elementary atoms can be accounted for by supposing that each atom has the power of fixing to itself a certain number of other atoms. It is as though a carbon atom, for example, had four arms by which it could grasp and be grasped by four hydrogen, or chlorine atoms, or two oxygen, or sulphur atoms, each of the latter having two arms of its own. It must, of course, be understood that this is a mere illustration, for no one believes that the atoms have real arms projecting from them. But the illustration gives a lively idea of the hypothesis; and symbols are sometimes written with arms (called bonds) proceeding from them. The following figures represent a few of these atoms and their compounds:

Monad atoms $\text{H}—$

Diad , $\text{O}—$

Triad ,, $\text{N} \equiv$

Hydrochloric Acid $\text{H}—\text{Cl}.$

Water $\text{O}—\text{H}.$
 $—\text{H}.$

Ammonia $\text{N}—\text{H}.$
 $—\text{H}.$
 $—\text{H}.$

Methane $\text{C} \begin{cases} \text{H} \\ \text{H} \\ \text{H} \\ \text{H} \end{cases}$

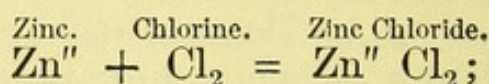
same bulk as one molecule of chlorine, but only that n molecules of hydrogen occupy, by their molecular movements, the same space as n molecules of chlorine.

2. Similarly, the hypothesis does not assert that the bulk of an atom of hydrogen is equal to the bulk of an atom of chlorine, but only that two atoms of hydrogen and two atoms of chlorine occupy, respec-

The hypothesis put in this form, appears somewhat fanciful, but it harmonizes well with known facts, and has the merit of assisting wonderfully in the comprehension of the complex compounds of organic chemistry.

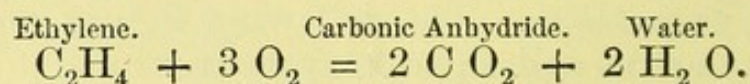
EQUATIONS, OR FORMULÆ OF CHEMICAL CHANGE.

Chemical changes of all kinds can be very conveniently represented in the form of equations, the formulæ for the substances concerned being written in the first half, and the formulæ for the new substances produced in the other. Thus the combination of zinc and chlorine is thus expressed :



which may be read in the following manner. One molecule of zinc added to one molecule (two atoms) of chlorine is equal to, or rather produces one molecule of zinc chloride.

In making out an equation of this kind, it is necessary, as with algebraical equations, to take care that every atom which appears in the first half shall duly appear and be accounted for in the second. The following equation, which expresses the oxidation of ethylene-gas, illustrates this :



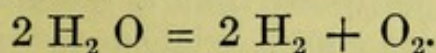
Two atoms of carbon, four of hydrogen, and six of oxygen are concerned in the change, and it will be seen that though differently arranged, they are all accounted for in the second half of the equation. The large figures in the above and in similar equations refer, it must be remembered, to the *whole molecule*. 2CO_2 , for instance, means two molecules of carbonic anhydride, containing together two atoms of carbon and four of oxygen. A few more equations, one or two of them rather complex ones,

tively, by their atomic movements, the space of one molecule. The atoms of matter may be almost infinitely small, as compared with the molecules, and now that we know approximately the size of the molecules we are no nearer to knowing that of the atoms.

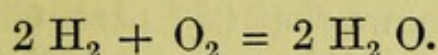
A molecule of hydrogen may, for aught we know, be comparable to an inflated bladder, with two peas rotating in it.

may, with advantage, be studied in this place. They represent chemical changes of several kinds.

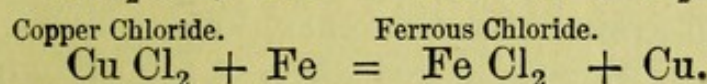
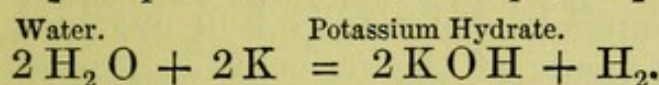
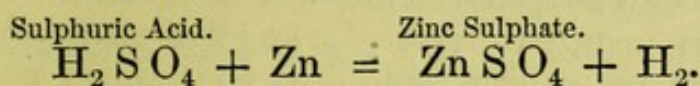
Decomposition of water :



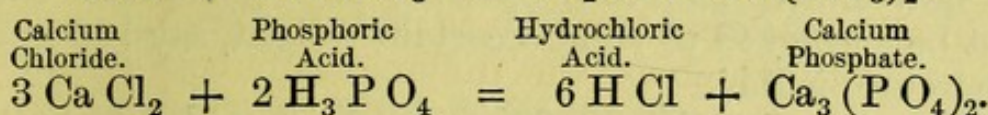
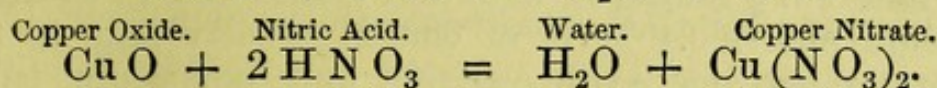
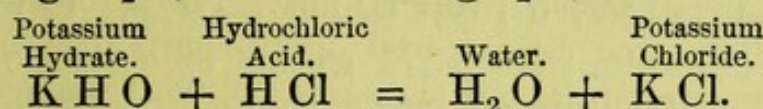
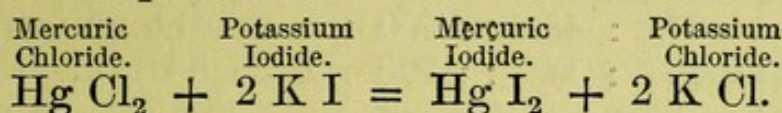
Synthesis of water :



Substitutions :



Double decompositions :



The brackets in the two last examples will readily be understood. $\text{Cu (NO}_3\text{)}_2$ expresses that one atom of the diad metal copper is united with two units of the monad radical NO_3 .

One great advantage of these equations is that they afford us the means of calculating the respective quantities by weight in which bodies act on one another. Take, for example, the above-described action of mercuric chloride and potassium iodide. HgCl_2 means one atom of mercury weighing 200 combined with two atoms of chlorine weighing $35.5 \times 2 = 71$, total 271. KI means one atom of potassium 39 and one atom of iodine 127, total 166; or, as 2KI is employed, 332. We therefore know that 271 parts (pounds, grammes, or tons) of mercuric chloride will act upon 332 parts of potassium iodide, and in a similar

manner we can calculate that 454 parts of mercuric iodide ($\text{HgI}_2 = \text{mercury, } 200 + \text{iodine, } 127 \times 2 = 254$), and 149 parts of potassium chloride will be produced during the change. Now suppose that a chemist has 100 grains of mercuric chloride, and wishes to know how much potassium iodide he must add to convert all the mercury into iodide. He knows that 271 parts of the chloride will require 332 of the iodide, and he has therefore only to perform a simple proportion sum.

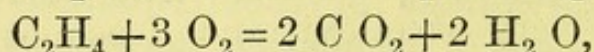
$$271 : 332 :: 100 : x = 122\frac{1}{2} \text{ grains.}$$

In like manner he can readily find out that he ought to obtain as the result of the action $167\frac{1}{2}$ grains of mercuric iodide and 55 grains of potassium chloride, for as 271 parts of mercuric chloride yield 454 parts of mercuric iodide and 149 parts of potassium chloride,

$$\begin{aligned} 271 : 454 :: 100 : x &= 167\frac{1}{2} \text{ grains} \\ \text{and } 271 : 149 :: 100 : x &= 55 \quad , , \end{aligned}$$

In short, to sum up the whole reaction, 100 grains of mercuric chloride, with $122\frac{1}{2}$ grains of potassium iodide, will yield $167\frac{1}{2}$ grains of mercuric iodide and 55 grains of potassium chloride. It is of course equally easy to find out how much mercuric chloride and potassium iodide must be employed to yield a certain weight, say 100 grains of mercuric iodide.

When the equations refer to gases they have another advantage. As every single formula for an element or compound denotes two volumes of gas, the volumes of different gases concerned in a reaction can at once be inferred from the equation. In the equation given above



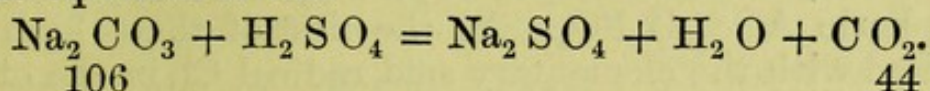
C_2H_4 means 2 volumes of ethylene, and 3O_2 means $3 \times 2 = 6$ volumes of oxygen; so that 2 volumes of ethylene require 6 volumes of oxygen for their complete oxidation. There will be produced during the action, 2CO_2 , that is, 4 volumes of carbonic anhydride and $2\text{H}_2\text{O}$, that is, 4 volumes of steam. For if CO_2 represents 2 volumes, 2CO_2 must evidently represent 4 volumes. It must not, however, be forgotten that these relations of volume are only correct if the pressure and temperature remain unaltered. If other-

wise, a correction must be made by the methods already given.

Taking the data already given, it is easy to calculate on the metrical system the absolute volume of any gas yielded by a certain weight of material, or the weight required to produce a certain volume of gas at standard pressure and temperature.

Examples.—1. What weight of carbonic anhydride gas can be obtained from 100 grammes of sodium carbonate?

The equation is this—

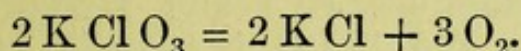


106 grammes of the carbonate yield 44 grammes of the gas. But as the formula for a gas always denotes 2 volumes, 44 grammes measure 22.4 litres. Hence

$$106 : 100 :: 22.4 : x = 21.1.$$

2. What weight of potassium chlorate must be decomposed if we wish to obtain 100 litres of oxygen?

Equation—



Here $2\text{KClO}_3 = 245$ grammes, and $3\text{O}_2 = 96$ grammes = 67.2 litres. Hence—

$$67.2 : 100 :: 245 : x = 364.6.$$

ACIDS, BASES, SALTS.

Allusion has already been made in Chapter II. to certain substances which have long been known under the respective names of *acids* and *bases*. Acids are sour and redden litmus, and bases, when soluble, have what is called an alkaline taste, and turn the colour of reddened litmus back again to blue. They have a kind of antagonistic function, and will *neutralize* one another. When an acid acts on a base, a new compound called a *salt* is produced, which commonly has no action on either blue or red litmus.

But although in well-marked cases acids, bases, and salts are so different from one another in properties, modern chemistry has taught us that the compounds usually known by those names bear so much resemblance to one another and to other compounds in *structure*, that it is impossible to frame accurate definitions for them. If possible

it would perhaps be as well to banish the terms altogether; but as that would produce great inconvenience, it is better to take such imperfect definitions as we can get.

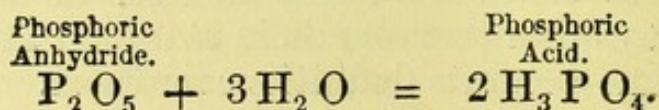
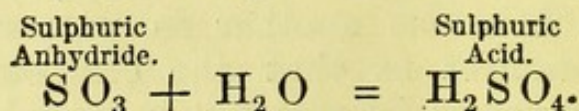
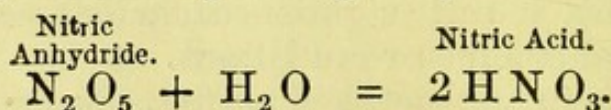
1. *Acids*.—An acid is composed of hydrogen with one of those atoms, or groups of atoms, which are called acid radicals. The hydrogen can be replaced by metals, in which case one of the compounds called salts is formed. Acids redden litmus, and are commonly sour.

The following are important examples:—Hydrochloric Acid, HCl ; Nitric Acid, HNO_3 ; Sulphuric Acid, H_2SO_4 ; Phosphoric Acid, H_3PO_4 . In the first of these hydrogen is united with the elementary acid radical Cl , in the others, with the compound radicals NO_3 , SO_4 , and PO_4 .

The imperfection of the definition will be apparent if we remember that we can only define an acid radical as a radical which, united with hydrogen, forms an acid.

Basity of Acids.—Acids are said to be *monobasic*, *dibasic*, or *tribasic*, according as they contain one, two, or three atoms of hydrogen which can be replaced by acids. In the previous examples HCl and HNO_3 are monobasic, H_2SO_4 is dibasic, and H_3PO_4 is tribasic. Of course this is equivalent to saying that the radicals of those acids are monad, diad, or triad. PO_4 is a triad radical, and its acid may be written in this way, $\text{H}_3(\text{PO}_4)'''$. Acids which contain more than one atom of replaceable hydrogen are said to be *polybasic*.

Oxygen Acids, or Oxy-acids.—Most acids contain oxygen as a part of their radical, and these are related in a very simple manner to a particular series of oxides, which, for that reason, are called *acid oxides* or *anhydrides*. When one of these anhydrides combines with water, an acid is formed, and when, on the other hand, water is removed from an acid, the corresponding anhydride is obtained.



One molecule of nitric anhydride, uniting with one of water, forms two molecules of nitric acid. In a similar manner, if from one molecule of sulphuric acid one molecule of water be taken, one molecule of sulphuric anhydride will remain.

It must not, however, be supposed the acids are always, or even generally, prepared in practice from the anhydrides. One anhydride, indeed, silicic anhydride, SiO_2 , will not combine directly with water, although its acid can be obtained by indirect means. The anhydrides of many acids have not yet been obtained, and a few, such as carbonic anhydride, CO_2 , are known to which no corresponding acids can be proved to exist. The formula for carbonic acid should be H_2CO_3 .

The acid anhydrides are now very commonly named as oxides. Thus SO_3 is called sulphuric trioxide, and CO_2 carbon dioxide.

2. *Salts*.—A salt is a compound containing a metal and an elementary or compound acid radical.

A salt only differs from an acid by containing a metal in place of hydrogen. Taking, for example, the salts of sodium, those corresponding to the acids already mentioned are—

Na Cl	Sodium chloride;	corresponding to	H Cl .
Na N O_3	„ nitrate	„	H N O_3 .
$\text{Na}_2 \text{S O}_4$	„ sulphate	„	$\text{H}_2 \text{S O}_4$.
$\text{Na}_3 \text{P O}_4$	„ phosphate	„	$\text{H}_3 \text{P O}_4$.

The proportion which the metal and radical bear to one another depends of course on the valency of each. Thus the chloride, nitrate, sulphate, and phosphate of the diad metal calcium, and of the triad metal bismuth, are formulated in this way: .

$\text{Ca}'' \text{Cl}_2$	$\text{Ca}'' (\text{N O}_3)_2$	$\text{Ca}'' (\text{S O}_4)''$	$\text{Ca}''_3 (\text{P O}_4)'''^2$.
$\text{Bi}''' \text{Cl}_3$	$\text{Bi}''' (\text{N O}_3)_3$	$\text{Bi}'''_2 (\text{S O}_4)'''_3$	$\text{Bi}''' (\text{P O}_4)'''$.

In many salts of polybasic acids the hydrogen is only partially replaced by metals. Thus we have—

Phosphoric Acid.	Sodium-dihydrogen Phosphate.	Disodium-hydrogen Phosphate.	Sodium phosphate, Trisodium Phosphate.
$\text{H}_3 \text{P O}_4$.	$\text{Na H}_2 \text{P O}_4$.	$\text{Na}_2 \text{H P O}_4$.	$\text{Na}_3 \text{P O}_4$.

It will be seen from the above remarks that acids are

really hydrogen salts. Some chemists indeed name them accordingly, and call H_2SO_4 "hydrogen sulphate," instead of sulphuric acid. The term acid is applied by them to the anhydrides, so that the formula for sulphuric acid becomes SO_3 . Many salts are known which are more or less irregular in their composition, but modern chemical theory enables us to give a tolerably satisfactory account of all of them.

3. *Bases*.—A base is a metallic hydrate or hydroxide (that is, a compound of a metal with the radical OH) which is capable of reacting with acids to form salts.

As OH is a monad radical, the constitution of hydrates is closely analogous to that of chlorides, nitrates, &c. For example :

Sodium Hydrate.

Na OH is like Na Cl and Na NO_3 .

Calcium Hydrate.

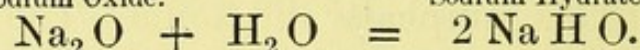
$\text{Ca}''(\text{OH})_2$ is like $\text{Ca}''\text{Cl}_2$ and $\text{Ca}''(\text{NO}_3)_2$.

Bismuth Hydrate.

$\text{Bi}'''(\text{OH})_3$ is like $\text{Bi}'''\text{Cl}_3$ and $\text{Bi}'''(\text{NO}_3)_3$.

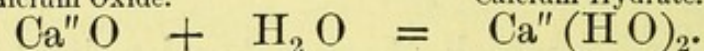
Bases, like acids, are related to a particular series of oxides called *basic oxides*. These oxides are sometimes called bases. They differ from the true bases by the elements of water.

Sodium Oxide.



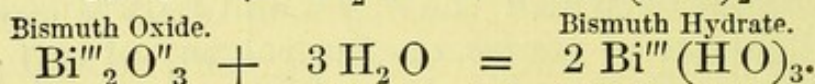
Sodium Hydrate.

Calcium Oxide.



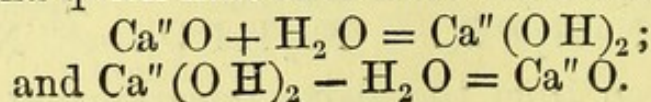
Calcium Hydrate.

Bismuth Oxide.



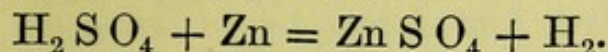
Bismuth Hydrate.

Some basic oxides, however (bismuth oxide, for instance), will not combine directly with water; and, on the other hand, some basic hydrates (sodium hydrate, for instance, cannot be dehydrated by heat. Calcium oxide and hydrate are examples of compounds which experience both changes with ease. The oxide (quick lime) combines eagerly with water, great heat is produced, and calcium hydrate (slaked lime) is formed. When slaked lime is heated, water is expelled, and quick lime once more obtained.

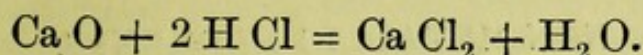


Formation of Salts.—Salts can be formed by a variety of processes, a few of which may be specified here, with one or two examples of each. It must not be supposed that every process is practicable in all cases.

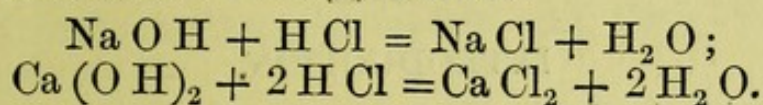
1. By the action of metals on acids :



2. By the action of metallic oxides on acids :

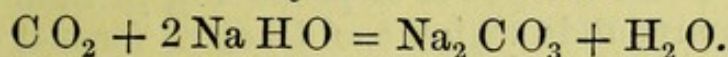


3. By the action of bases on acids :

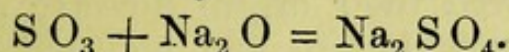


This is the case before referred to.

4. By the action of anhydrides on bases :



5. By the action of anhydrides on metallic oxides :



It will be seen that in most of these reactions water is formed simultaneously with the salt.

THE LAW OF NEWLANDS, OR THE PERIODIC LAW.

If the elements are arranged in the numerical order of their atomic weights, it will be seen that, omitting hydrogen, the lower members arrange themselves in *octaves* or groups of seven, and that starting from any one of these elements, the eighth is in many physical and chemical respects analogous to it.

H 1. Li 7. Be 9. B 11. C 12. N 14. O 16. F 19.

Na 23. Mg 24. Al 27. Si 28. P 31. S 32. Cl 35.5.

K 39. Ca 40. Sc 44. Ti 48. V 51.

Allowing for unknown elements, the table may be made comprehensive, and the law is of daily increasing importance in chemistry.

PART II.

NON-METALLIC ELEMENTS.

INTRODUCTION.

THE old distinction between organic and inorganic chemistry has faded away. It was formerly believed that vegetables and animals had the power of producing in their organisms chemical compounds which could not be formed artificially in the laboratory. Vast numbers of compounds were known which owed their origin directly or indirectly to the animal or vegetable kingdom, and which could not be obtained from any other source. To the department of chemistry which dealt with such compounds, the name "Organic Chemistry" was very properly applied. But the progress of scientific research has taught us to manufacture a great number of these compounds from inorganic materials, and there is therefore no longer any reason why they should be separated from other compounds in a general system of classification.

The so-called organic compounds have, however, one feature in common. They all contain carbon, and it is therefore convenient for purposes of study to retain them in a separate department of chemistry. Moreover, although some few carbon compounds enter into the composition of important minerals, a very large number of them bear some direct or indirect relation to the processes of life, and the term "organic" may therefore, in this limited sense, still be applied to them. It is only necessary to bear in mind that what is still generally called organic chemistry is but the chemistry of carbon compounds. In the present division of the book we shall say as little about it as possible.

A less important distinction is that which is commonly made between metallic and non-metallic elements. No exact definition can be attached to the word metal, no one property can be connected exclusively with it, and there is no line of demarcation that can be drawn between the two classes. We shall trouble ourselves very little with the precise meaning of the word metal, but shall use it in its current sense, and shall introduce the elements in the order which appears most simple and convenient, grouping them according to their valencies. The present part contains an account of the elements generally classed as non-metallic. Part III. is devoted to the metals, and Part IV. to organic chemistry.

CHAPTER I.

NON-METALLIC MONADS.

(Hydrogen, Chlorine, Bromine, Iodine, Fluorine.)

HYDROGEN.

Symbol, $H = 1$. Formula, H_2 .

THIS element in its chemical relations resembles the monad metals. It stands alone among the non-metals, and it must, for convenience, be studied first.

PREPARATION.

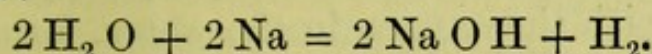
Hydrogen is a gas which can only be liquefied by intense cold. It occurs in nature almost exclusively in a state of combination, but various means are known by which it can be set free and obtained in a state of purity. We have already seen (p. 12) that when a current of electricity from a tolerably powerful battery is transmitted through water, acidulated with sulphuric acid, the water is decomposed, oxygen goes to one pole of the battery and hydrogen to the other; and both gases can easily be collected.

Electrolysis of Water.—Other compounds containing hydrogen yield it up under similar treatment.

The ordinary processes for preparing hydrogen almost all involve the use of some metal. A certain number of the metals have the power of displacing hydrogen from its combinations, either easily or with difficulty. For common purposes zinc or iron is used.

Experiment 1.—Boil some water for fifteen minutes, that all the air contained in it may be expelled; let it cool, and fill a bowl and a test-tube with it; close the latter with the finger, and invert the mouth of it under the

water in the bowl. Now wrap in a scrap of blotting-paper, a piece of sodium of the size of a small pea, and thrust it quickly under the mouth of the test-tube; the metal being lighter than water, ascends into the tube, floating there with a rotatory motion; a gas is evolved from the water, and collects in the upper part of the tube. This gas is hydrogen. The metal sodium displaces one half of the hydrogen from water, in the manner shown in the following formula:



The compound NaOH is called sodium hydrate. It remains dissolved in the water, and may be obtained in a pure state if the water is evaporated off. Close the tube again with the finger, remove it from the bowl, and apply a light to the mouth, the gas will burn with a flash of light. This experiment proves hydrogen to be a *combustible* gas. Pour into the bowl some solution of litmus, which has been reddened with a drop of vinegar or other acid, the litmus will be immediately changed to blue, showing that sodium hydrate is one of the substances called *bases*.

Experiment 2.—Lay a piece of blotting-paper on the surface of some water contained in a saucer, and throw upon it a small piece of sodium; an energetic decomposition of the water will take place, and in a few moments the sodium will apparently burst into flame, and burn for some time with an intense yellow colour. This apparent combustion of the sodium is really due to the burning of the hydrogen, set free by that metal, which is inflamed by the intense heat which accompanies its evolution. This experiment differs only from the preceding one inasmuch as in the former case the hydrogen is collected, while in the latter it is burnt as it is liberated. The sodium hydrate may be rendered evident as before by the addition of red litmus solution to the water.

Experiment 3.—Throw a piece of potassium on some water (the blotting-paper may be dispensed with in this case), the same effect as in the preceding experiment will occur, the water will be more violently decomposed, and the hydrogen—but apparently the potassium—will burn

instantaneously with a beautiful violet flame. Potassium hydrate, KOH, remains in solution in the water. It is a base like sodium hydrate.

Experiment 4.—What sodium and potassium accomplish at ordinary temperatures, iron can do if it be heated to redness. Pass water in the form of steam, obtained by boiling the water in a flask, or a retort, through a red-hot iron pipe, as a gun or gas barrel, containing small iron nails or wire. At this high temperature the iron in the pipe unites with the oxygen in the water, forming a black oxide of iron, and the hydrogen is set free and may be collected in the manner described below. The reaction is as follows :

$$4 \text{H}_2\text{O} + 3 \text{Fe} = \text{Fe}_3\text{O}_4 + 4 \text{H}_2.$$

The best means, however, for obtaining hydrogen is by the decomposition of an acid by iron or some other metal, zinc being generally chosen for the purpose.

Collection of Gases.—For collecting considerable quantities of gases the following contrivance, called a *pneumatic trough*, may be used. Make a shelf out of slate or a piece of lead, from three to four inches broad, and so long that it will rest about half-way up the sloping sides of a pan or wash-hand basin; cut a hole about half an inch in diameter through the centre of the shelf, and having placed the latter in position, pour into the pan sufficient water to cover it an inch deep. The shelf is for the purpose of supporting the vessel intended for the reception of the

Fig. 37.



gas, which, filled with water, is placed with its mouth exactly over the hole. The gas to be collected is then delivered from a tube, the extremity of which is placed directly under the mouth of the inverted vessel. The accompanying figure conveys an idea of the kind of apparatus required. A rectangular box with a shelf at one end, about two inches from the top, is more convenient.

Experiment 5.—Put half an ounce of *granulated zinc** in

* Zinc is granulated by being melted in an iron ladle, and poured from the height of a few feet into cold water.

a bottle or flask, and pour over it a little water. No action takes place, but if a small quantity of sulphuric acid be gradually added, effervescence, and heating of the mixture will ensue. The effervescence is caused by the escape of *hydrogen-gas*. Insert into the mouth of the bottle a cork, which has previously been perforated and fitted with a tube bent so that the mouth of it may be conveniently placed beneath the hole in the shelf of the pneumatic trough.* Allow time for the hydrogen to displace the air contained in the flask and bent tube (two or three minutes is sufficient if the effervescence is tolerably brisk). While the air is being expelled, preparation may be made for collecting the gas by filling several bottles quite full of water. One of the bottles is then closed with a smooth card or glass plate, and rapidly inverted in the pneumatic trough with its mouth directly over the hole in the shelf. The tube being then placed beneath, the gas will ascend and displace the water in the bottle. When the entire displacement is effected the bottle is corked or stoppered while still in the trough, removed and replaced by another bottle, inverted in the same manner, and so on until the evolution of gas ceases.

There is one indispensable caution to be observed in experimenting with hydrogen, which is, *not to begin to collect the gas until all the atmospheric air existing in the flask has been expelled*, as otherwise an explosion might take place. To be quite safe, it is as well to reject the first bottleful of hydrogen collected.

PROPERTIES.

Experiment 6.—Inflame hydrogen contained in a bottle, and immediately pour in some water. The water does not extinguish the flame, but rather increases it, since it rapidly forces the gas out of the flask. The gas does not burn in

Fig. 38.



* Instead of a rigid glass tube, it is more convenient for this and similar purposes to carry the gas to the pneumatic trough by means of a piece of india-rubber tubing, of about one-eighth of an inch internal diameter. It may be slipped over a short piece of glass tubing which passes through the cork.

the interior of the vessel, but only on the outside, where it is surrounded by atmospheric air.

Experiment 7.—Open a bottle of hydrogen under an inverted tumbler, and, after a minute or two, apply a lighted taper to the mouth of the tumbler. A flame will burst forth from the tumbler with a slight report. The gas has ascended from the bottle into the tumbler, and is consequently *lighter* than common air. In this experiment the lower vessel must not be immediately exposed to the lighted taper, because, if all the hydrogen is not displaced, an explosion might ensue that might break the bottle; but if the taper be applied after a few minutes have elapsed, the bottle will be found no longer to contain any combustible gas, the gas having entirely escaped.

Hydrogen is the *lightest* of all gases. Its specific gravity is 1, and $14\frac{1}{2}$ measures of it weigh only as much as one measure of atmospheric air. On account of this lightness, it may be used for filling balloons.

Experiment 8.—If, instead of the glass tube, a piece of tobacco-pipe or a short piece of glass tube drawn out at one end to a jet be adapted to the cork of the flask from which hydrogen was evolved, and the gas then lighted, it will burn like a taper.

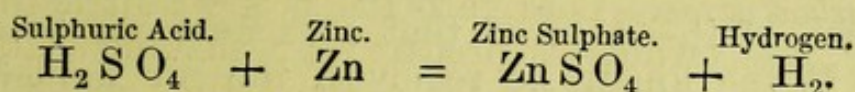


To kindle the gas, instead of a match or a taper, very finely divided platinum may be employed. This can be prepared in a few minutes by dropping a solution of platinum chloride on blotting-paper, attaching it to a wire, and igniting it over a spirit-lamp, till nothing but a grey coherent ash remains. The platinum is thus reduced to an extremely minute state of subdivision, and in this state it exhibits the remarkable property of igniting in hydrogen and inflaming it. It is called spongy platinum, and is employed as tinder in the well-known *Döhbereiner's lamp*. A sheet of "platinum leaf," prepared like gold leaf, is excellent for the experiment. It is well to heat it to redness first.

Experiment 9.—To observe the remarkable lightness of hydrogen, a small balloon of gold-beater's skin may be filled with hydrogen by means of the apparatus used in

Experiment 5. The balloon is squeezed flat to expel the air, and the tube delivering the gas is passed a short way into its orifice, and secured there with a piece of thread. When the inflation is complete, the tube is withdrawn, and the twine tightened simultaneously, thus preventing any escape of gas. The balloon, if unimpeded, will then ascend to a great height. It may be made captive with a piece of thread.

Experiment 10.—Pour the contents of the flask in which the hydrogen was generated (Experiment 5) into a porcelain dish, boil until they are reduced in bulk to one-half or thereabouts, and filter them. A black residue will remain on the filter, which consists of the impurities contained in the zinc; the zinc itself has been dissolved, and has been converted into a *salt* called zinc sulphate, which, on the cooling of the solution, is deposited in colourless crystals. The reaction which takes place between zinc and sulphuric acid is represented by the following equation :



CHLORINE.

Symbol, Cl = 35.5. Formula, Cl₂.

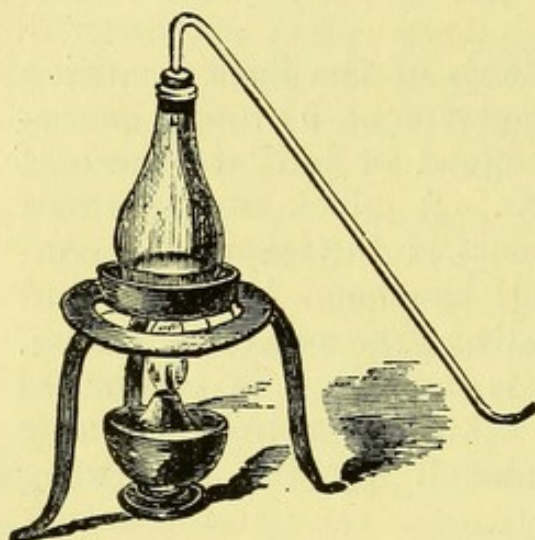
Chlorine is one of a group, the members of which are characterized by a remarkable similarity of chemical properties. It consists of the elements chlorine, bromine, iodine, and fluorine, the three first of which are often termed *halogens* (from *ἅλς*, sea-salt), in allusion to their marine origin.

Chlorine is a greenish-yellow gas, $2\frac{1}{2}$ times heavier than air, suffocating and irrespirable unless very much diluted with air. Its odour, when very dilute, is somewhat like that of sea-weed, a peculiarity which it shares with the other halogens. Chlorine occurs only in combination; chiefly with sodium as common salt (sodium chloride, NaCl), which forms immense deposits in England and elsewhere, and is the chief solid ingredient of seawater.

PREPARATION.

Experiment 1.—Pour one ounce and a half of hydrochloric acid upon a quarter of an ounce of finely-powdered black oxide of manganese, and heat it gradually in a flask, to which is adapted a bent glass tube; a yellowish-green gas

Fig. 40.



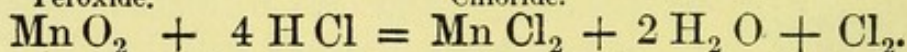
is disengaged, which is collected by the process already described. The pneumatic trough is, however, filled with warm water instead of cold. This gas is *chlorine* (from *χλωρός*, *green*). Fill with it several six-ounce bottles of white glass, and cork them up. Fill, likewise, a bottle with two-thirds of chlorine and one-third of water, and shake it up; suction is exerted upon a finger which closes

the mouth of it,—a proof that a vacuum has been produced. If the finger be removed, the air immediately rushes in. This vacuum was caused by the chlorine having dissolved in the water, which might be inferred also from the disappearance of the yellow colour from the upper part of the bottle. One measure of cold water dissolves two measures of chlorine. This solution is called *chlorine water*.

The mode in which chlorine is formed in this experiment is shown in the following formula :

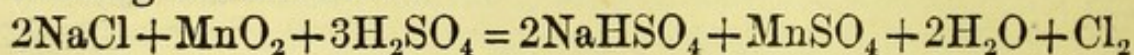
Manganese
Peroxide.

Manganese
Chloride.



When the evolution of gas has quite ceased the liquid in the flask may be filtered and evaporated, when it will yield on cooling pink crystals of MnCl_2 .

Experiment 2.—Chlorine may also be prepared from common salt by mixing three quarters of an ounce of it with half an ounce of black oxide of manganese, two ounces of sulphuric acid, and one ounce of water, and heating the mixture :



Chlorine acts as a *poison* on being inhaled; hence, *care must be taken not to inhale it* while preparing it. For greater security, pour some drops of alcohol and ammonia upon a cloth and wave it frequently in the air; the chlorine contained in the air will then be so altered that it will lose its injurious properties.

PROPERTIES.

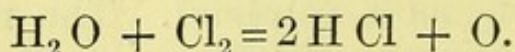
Experiment 3.—In order to recognise the odour of chlorine, smell chlorine water (but not the gas) cautiously; the chlorine water may be tasted also without danger.

Experiment 4.—If a flask containing chlorine gas be exposed to the air for a short time, no diminution of the chlorine will be perceptible; but if the flask be inverted it will soon contain only atmospheric air. Chlorine is two and a half times *heavier* than common air, and may be easily poured from one vessel to another like water without material waste; its specific gravity is 35.5 ($H = 1$).

Experiment 5.—Introduce a piece of litmus-paper into chlorine gas, and it becomes white; pour chlorine water upon red wine, and ink, and both the liquids will lose their colour. *Chlorine bleaches and destroys most colours derived from the animal or vegetable kingdom.* In consequence of this property, chlorine has become a most important agent in bleaching; and linen, cotton, paper, and other materials, may be rendered perfectly white by it in a few hours; while, by the old method of laying them on the grass in the sun, weeks, and even months, were required for effecting it. Substances called *antichlors* are sometimes used to remove the last traces of chlorine. Sodium thiosulphate is the most powerful. The modern method of bleaching is very excellent, and does not injure the strength of the fabric, provided all the chlorine be completely removed after the bleaching is finished, which is not so easily done as might be supposed. If this precaution be not observed, or if the chlorine water be too strong or in excess, then, indeed, after the colour is destroyed, the fibres of the yarn or fabric itself will be attacked. The substance commonly called *chloride of lime* is now used instead of chlorine. It is a salt from which chlorine is easily disengaged, even by mere exposure to the air.

Experiment 6.—Apply chlorine water to decaying and nauseous substances (water in which flowers have been kept, manure, rotten eggs, &c.); the bad odour will at once entirely vanish. Thus it not only *decomposes* colours, but also the *volatile combinations* formed during putrefaction, and which occasion disagreeable odours. It acts in a similar manner also upon morbid matter (matters of contagion, miasmata), which, being diffused in the air or attached to clothes and beds, may communicate disease. Chlorine is, therefore, a powerful *disinfecting agent*, and is used for purifying putrefying matter and infected atmospheres. Musty casks may also be purified by washing them first with chlorine water, and then with some milk of lime. Mouldy cellars, in which milk or beer cannot be kept without turning sour, are again rendered serviceable for a long time by fumigating them with chlorine-gas, or by washing them with chlorine water, or a solution of chloride of lime.

Experiment 7.—Fill a small bottle with chlorine water, and invert it in a vessel filled with water; if this is put away in a dark place, it remains unchanged; but if it is exposed to the sun, a colourless gas will collect in the upper part of the flask, in which a glowing taper will inflame; this gas is oxygen. After some days the water will entirely lose its odour of chlorine, and will have acquired a sour taste, and instead of bleaching blue litmus-paper, it will redden it. Three elements only were present, the constituents of water and chlorine; thus it is obvious that the chlorine must have united with the hydrogen of the water to form hydrochloric acid, the oxygen being set free:



Bottles in which chlorine water is kept should, therefore, be protected from the light, and this can be most conveniently done by pasting black paper round them.

The bleaching and disinfecting power of chlorine is now easily explained by its strong affinity for hydrogen. All animal and vegetable substances contain hydrogen, which is taken from them by chlorine. By the abstraction of the hydrogen, the colouring matter becomes colourless, the odorous principles scentless, the morbid matter in-

noxious, insoluble substances are very frequently rendered soluble, &c.

Experiment 8.—Provide two bottles of chlorine, and place in one of them some dry calcium chloride. This salt eagerly absorbs water, and will thoroughly dry any gas in contact with it. In a few hours introduce a piece of blue litmus-paper into the dry chlorine; no change in the litmus will be apparent, or, at most, it will be but slightly reddened. Perform the same operation in the other bottle, and the litmus-paper will be rapidly bleached; proving that the presence of water is necessary to enable chlorine to exert its power in this respect.

Experiment 9.—Put into chlorine water some gold-leaf; it will soon disappear, as the element chlorine combines with the element gold. The combination is called *auric chloride*; it is soluble in water. Chlorine has a very *great tendency to combine with the metals*. These combinations comport themselves as salts; they are called *metallic chlorides*, and most of them are soluble in water.

Experiment 10.—Pour into a vessel filled with chlorine-gas a little metallic antimony, in fine powder; it will fall in a red-hot state to the bottom, as though it were a shower of fire. The red heat is caused by the violent combination of the chlorine with the antimony. The white smoke which fills the flask is the new combination formed, viz., *antimonic chloride*. If a fine brass wire, on which a piece of tinsel (Dutch gold) has been fastened, be introduced into chlorine-gas, both will burn with vivid combustion, and with the emission of fumes. Here combustion is another name for combination with chlorine. Brass consists of zinc and copper: accordingly, chlorides of zinc and copper are formed. Both dissolve in water, and the copper chloride imparts to the solution a green tinge.

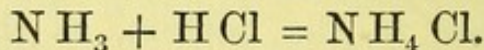
Experiment 11.—If a piece of sodium of the size of a pea is thrown into a cup containing chlorine water, it will move rapidly round, just as in common water, with a hissing noise, and finally disappear; but if a sufficient quantity of the chlorine was present, the liquid will not afterwards give a basic reaction, as in Experiments 1 and 2, *Hydrogen*; neither will it have an alkaline, but a saline

taste. If allowed to evaporate gradually over a warm stove, small cubic crystals remain behind, the constituents of which are chlorine and sodium. Thus, from these two elements a *salt* has been formed, familiarly known as *common salt*, NaCl.

Chlorine, like oxygen and sulphur, often unites in several proportions with a substance. Thus, there are two different chlorides of mercury.

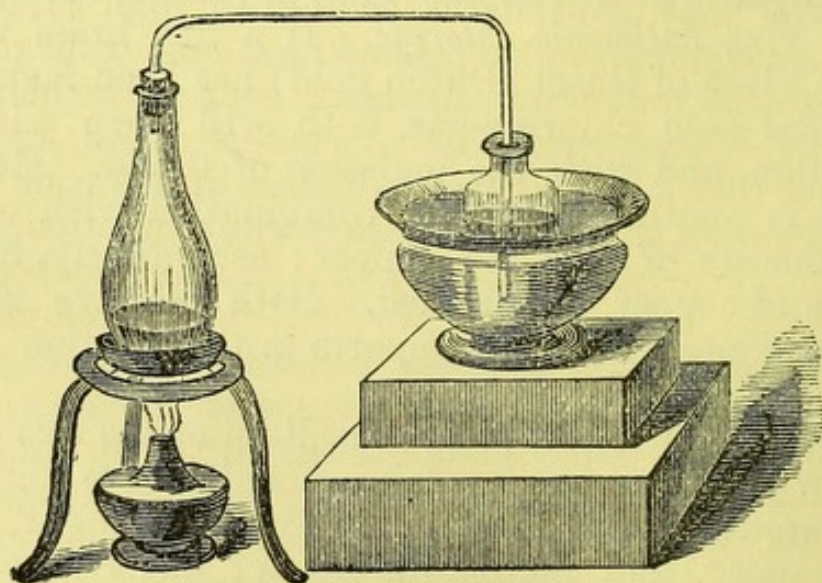
HYDROCHLORIC ACID, HCl.

Experiment 1.—Put into a porcelain capsule a few grains of common salt, and pour a little sulphuric acid upon it; there escapes, with effervescence, a gas, which has an acid taste, fumes in the air, and reddens moistened blue test paper; this gas is *muriatic acid*, or *hydrochloric acid*. If you pour some ammonia upon a shaving, or slip of paper, and wave the latter to and fro over the capsule, a thick white smoke is formed; and the odour of both the hydrochloric acid, and also the pungent fumes of the ammonia, vanish. The two gases combine and form solid *ammonium chloride*:



Experiment 2.—Mix carefully in a flask a quarter of an

Fig. 41.



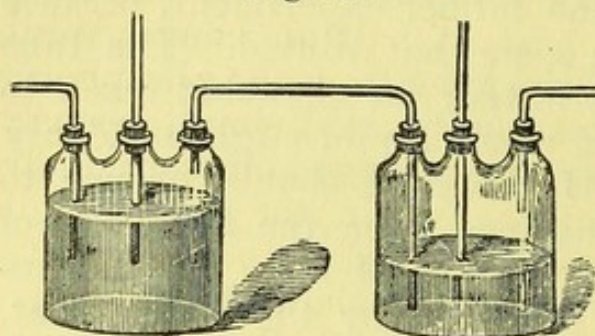
ounce of water with three-quarters of an ounce of sulphuric acid, and after the mixture has become cold, add to it half an ounce of common salt. Adapt to the neck of the flask

a cork provided with a glass tube, the long limb of which passes into a phial, containing one ounce of water. If you heat the flask in a sand-bath, the hydrochloric acid escapes, but more quietly than in the former experiment, because the sulphuric acid has been somewhat diluted. The tube must only just dip into the water; for should it reach to the bottom of the phial, the whole liquid might suddenly flow back into the flask, if the heat should chance to slacken, as it might, for instance, from the flickering of the lamp by an accidental current of air. The hydrochloric acid is so eagerly absorbed by the water, that, when the evolution of the gas diminishes, a vacuum is formed in the tube and flask; the pressure of the exterior air then forces the water up into the flask. When a gaseous body condenses into a liquid, it no longer requires the latent heat by which it became gas or vapour, and therefore this heat is set free. From this it follows that the water in which the acid condenses or dissolves must soon become warm. But warm water takes up much less gas than cold; accordingly, in order to obtain a concentrated solution of hydrochloric acid gas, we must place the phial in a basin of cold water. When the liquid in the receiver has sufficiently increased, one of the blocks must be withdrawn from beneath, so as to keep the end of the tube near the surface of the liquid. The solution thus obtained has an intensely acid taste and reaction; it is called *hydrochloric acid*, but is often known by the name of *muriatic acid*. One measure of cold water absorbs more than four hundred measures of hydrochloric acid gas; the strong hydrochloric acid thus obtained fumes in the air, because a part of the gas escapes. If you heat it to boiling then half of it escapes, and an acid only half as strong remains behind; but this is always somewhat *heavier* than water.

The hydrochloric acid of commerce is commonly yellow, and contaminated with sulphurous acid, sulphuric acid, chlorine, iron, and arsenic. It is manufactured from common salt and sulphuric acid; but, instead of glass vessels, large iron retorts are employed, capable of containing some hundredweights of common salt. The gas is conducted into several bottles or jars filled with water,

and connected with each other. When the water in the first vessel becomes saturated with the hydrochloric acid

Fig. 42.

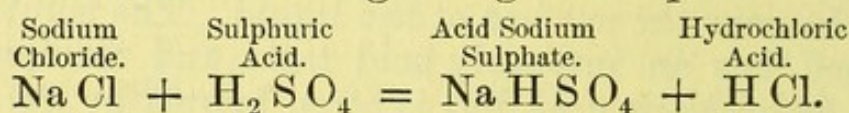


gas, the gas passes over into the second, then into the third vessel, and so on, saturating each successively. This is a very convenient method of conducting gases through liquids. Such vessels, which are commonly provided with

two or three necks, are called *Woulf's bottles*. The upright tube in the middle neck serves as a safety tube, that is, it prevents the liquid from being forced back; if a vacuum is formed in one of the bottles, the air enters through this tube.

Experiment 3.—If instead of dipping into water, the delivery-tube is passed to the bottom of a dry bottle, the heavy gas will displace the air, and fill the bottle. Its solubility may then be shown by inverting the bottle quickly in water. The water will rush up and fill the bottle.

Common salt consists of chlorine and sodium. With sulphuric acid, the following change takes place:



The constituents of hydrochloric acid gas are equal atoms of chlorine and hydrogen, and it is represented by the formula HCl .

If you fill a glass vessel half with chlorine and half with hydrogen, and put it in a dark place, no union ensues; but it takes place instantaneously when a light is applied or when the vessel is exposed to the direct rays of the sun. The union is accompanied by a violent detonation, which often breaks the vessel, so that it is not advisable to perform this experiment except on a very small scale. In the diffused light of day the combination takes place slowly.

Experiment 4.—Put some iron nails into a phial, and pour upon them some hydrochloric acid; brisk effervescence

will ensue. When this has continued some minutes, hold a burning taper over the mouth of the phial; the gas which escapes takes fire; it is hydrogen. The acid is decomposed, and its second constituent, chlorine, combines with the iron. The iron disappears, and is dissolved; that is, it combines with the chlorine, forming a soluble compound. When the effervescence has ceased, heat the phial by placing it in hot water, and afterwards pour its contents upon a filter of white blotting-paper. Put the liquid which passes through (*the filtrate*) in a cool place; a salt is deposited from it in greenish crystals, called *ferrous chloride*, FeCl_2 .

Many other metals are also dissolved, like iron, by hydrochloric acid, and converted into chlorides.

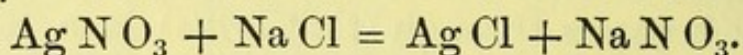
Experiment 5.—Pour some hydrochloric acid upon iron-rust that has been put into a test-tube; it dissolves, but without evolution of gas. In this case, the hydrogen of the hydrochloric acid meets with a body with which it can combine, namely, the oxygen of the oxide or rust of iron; and it does combine with it, forming water. The yellowish-brown solution, which is difficult to crystallize, yields, upon evaporation, a brown mass called *ferric chloride*, Fe_2Cl_6 . This salt contains one half more chlorine than the former. Hydrochloric acid is very often used for dissolving metallic oxides.

Experiment 6.—Dissolve some of the ferrous chloride, obtained according to Experiment 4, in a little cold water, and then add some chlorine water; the greenish colour is converted into a yellow colour, and the solution yields, on evaporation, brown *ferric chloride*. The chlorine converts the *ferrous* chloride into *ferric* chloride.

Experiment 7.—Dissolve some carbonate of soda (sodium carbonate) in water; the solution turns red test-paper blue; it has a basic reaction. Drop carefully into the solution some hydrochloric acid, until neither the red nor the blue paper is affected by it. If you put the liquid in a warm place a salt will be deposited in small cubes; you readily perceive, both by the shape of the crystals and by the taste, that it is common salt. Carbonic anhydride escapes with effervescence.

Experiment 8.—To a little water in a test-tube add a

drop of hydrochloric acid or a grain of a soluble chloride, such as common salt, and then a few drops of a solution of silver nitrate (lunar caustic); a white cloudiness is formed which does not happen in pure water. This cloudiness or *precipitate* is due to the formation of silver chloride, which is insoluble in water, but it may be dissolved by adding some *solution of ammonia*. Nitrate of silver is a most accurate test for hydrochloric acid or for chlorides:

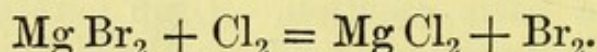


BROMINE.

Symbol, Br = 80. Formula, Br₂.

Bromine is a deep brownish-red, heavy, and very volatile liquid. Its name is derived from the Greek word *βρῶμος*, signifying a disagreeable odour. Bromine, at common temperatures, emits yellowish-red fumes, which have a penetrating and offensive odour, resembling that of chlorine. It produces a yellow colour with starch, is sparingly soluble in water, but very soluble in ether. Its specific gravity is 3.

Bromine occurs in many mineral waters, some of which, as the Kreutznach spring, are comparatively rich in it, and form the source from whence much of it is obtained on the Continent. A very considerable supply is also derived from sea-water. A large quantity of common salt is obtained by evaporating sea-water, and it is from the residual liquid, which is technically called *bittern*, that bromine is obtained. The *bittern* is mixed in a retort with manganese peroxide, and hydrochloric acid. On the application of heat the chlorine liberated from the acid decomposes the magnesium bromide contained in the *bittern*, and bromine, mixed with water, distils over:



The bromine may be separated from the water by shaking it with ether or chloroform. The ethereal solution, which has a brown colour, floats on the water and can be drawn off. The bromine is then extracted by a somewhat complex process.

Experiment 1.—The processes for the preparation of

bromine may be imitated by the student by dissolving a few grains of potassium bromide in a little water contained in a long test-tube; this may represent the bittern. A little chlorine water will liberate the bromine, and the solution will become brown. If a drachm or two of ether is then added and the mixture violently shaken for a minute, the ether will remove all the bromine, and on standing a little while will float as a dark layer on the surface of the water. Pour off the dark-coloured ether into a porcelain basin, and add a few drops of potassium hydrate (caustic potash); it will become colourless. Heat it gently until all the liquid is driven off and a white substance will remain behind, which on cooling may be mixed with a few grains of manganese peroxide, and two or three drops of sulphuric acid, and again heated. Free bromine will then be evolved as brown vapours, very readily seen against the white sides of the basin.

Experiment 2.—Pour a few drops of bromine into a small phial with some distilled water, and shake the mixture well. The bromine will partly dissolve and form a brown solution. This may be preserved and labelled “Bromine solution.”

Experiment 3.—Add a few drops of the above solution to a little very weak mucilage of starch, made by mixing a few grains of starch with a *little* cold, and then adding a considerable quantity of boiling water, with continual stirring. A pale yellow colour will be produced, owing to the formation of bromide of starch.

Experiment 4.—To a drop of bromine contained in a wine-glass add a *minute* fragment of phosphorus, which must be cut off under water, as it is very inflammable, and carefully and quickly dried in a piece of blotting-paper. The best plan is to drop the phosphorus from the point of a knife on to the bromine. The phosphorus will be immediately inflamed, and a smart explosion will occur.

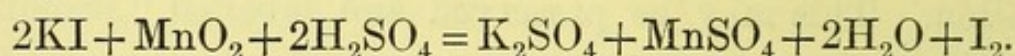
IODINE.

Symbol, I = 127. Formula, I_2 .

Iodine is a black solid of metallic lustre; it smells somewhat like chlorine, has a pungent taste, and stains the skin brown.

It occurs in small quantity in sea-water, from which it is separated and assimilated by the various sea-weeds, which thus accumulate a considerable quantity; and it is from the ashes of certain kinds of sea-weeds that iodine is always obtained, for sea-water contains much too small a quantity to render its extraction profitable. The sea-weed is burnt and the ash, or *kelp*, as it is called, contains the iodine in the forms of magnesium iodide, MgI_2 , and sodium iodide, NaI , and also corresponding bromides. The kelp, after some preliminary treatment, is heated with black oxide of manganese and sulphuric acid, and yields iodine and bromine, just as the chlorides yield chlorine when treated in the same way.

Experiment 1.—To a few grains of potassium iodide, or some other iodide, in a test-tube, add an equal weight of black oxide of manganese, and a drop or two of sulphuric acid, and heat gently. Beautiful violet vapours of iodine will be produced, and will condense on the sides of the test-tube as black, shining spangles. The following change occurs:



Experiment 2.—Dissolve a grain or two of potassium iodide in a little distilled water in a test-tube, and add a little solution of silver nitrate. A light yellow precipitate (silver iodide) will be thrown down, which will not be re-dissolved on the addition of solution of ammonia, thus clearly distinguishing it from the similar white precipitate obtained with chlorides.

Experiment 3.—Put 24 grains of iodine into a flask, and pour over them half an ounce of strong alcohol; if the iodine is pure it will entirely *dissolve*. This dark brown solution is called *tincture of iodine*. Water dissolves only a trace of iodine, but yet is rendered yellow by it.

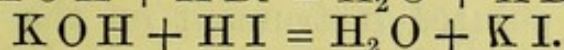
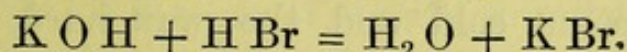
Experiment 4.—Put a little iodine upon a knife, and hold it over the flame of a lamp; the iodine melts, and is afterwards converted into a *violet-coloured gas—vapour of iodine*. As the iodine fumes are nearly nine times heavier than common air, they sink in it. Iodine owes its name to the colour of its vapour, the Greek word $\iota\omega\delta\eta\varsigma$ meaning *violet coloured*. The vapour appears more beautiful

when the iodine is heated in a small flask. After cooling, the walls of the flask become lined with small brilliant crystals of solid iodine, affording an example that regular crystals may be formed when bodies pass from the aeriform into the solid state.

Experiment 5.—Boil one grain of starch in a test-tube with one drachm of water, and after cooling add to the thin paste thus obtained a few drops of tincture of iodine ; *the iodine combines with the starch ; the combination is of a deep blue colour.* The blue colour disappears on careful heating, but returns again on cooling. If one drop of the starch paste is mixed with one quart of water, even at this extreme dilution, the iodine tincture will impart to it a violet tinge. Consequently it is an exceedingly delicate test for starch, and starch, on the other hand, for iodine. If a little tincture of iodine is dropped upon flour, bread, potatoes, &c., the presence of starch in these substances will at once be indicated.

COMPOUNDS OF BROMINE AND IODINE WITH HYDROGEN.

Hydrobromic Acid, HBr, and Hydriodic Acid, HI.—Both of these acids closely resemble hydrochloric acid. They are colourless gases, which fume strongly in air, and are readily soluble in water, forming acid solutions. With bases they form iodides and bromides :



Both bromine and iodine are faithful companions of chlorine ; wherever common salt occurs, whether in the earth, the sea, or mineral springs, small quantities of them are present, not in a free state, however, but combined with metals. The different sea-weeds attract these combinations from the sea-water, and from these sea-weeds iodine and bromine are extracted. Both have poisonous properties.

FLUORINE.

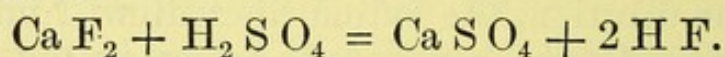
Symbol, F = 19. Formula unknown.

Fluorine in the free state is but little known. When liberated from its compounds it acts energetically on all

substances of which vessels are ordinarily made. The mineral known as fluor-spar consists of fluorine and calcium.

Hydrofluoric Acid, HF.

Experiment 1.—Fashion a small basin out of a piece of thin sheet lead, place in it about a quarter of an ounce of powdered fluor-spar, and half an ounce of strong sulphuric acid, and apply a gentle heat. Fumes of hydrofluoric acid gas will arise, which are extremely irritating to the eyes, and which resemble those of hydrochloric acid. Fluor-spar (calcium fluoride) is decomposed by sulphuric acid just as common salt is :



Hydrofluoric acid is intensely corrosive, and it attacks and destroys glass and earthenware and most other materials. But it does not act upon lead; therefore vessels made of that material must be used to prepare it in. The most interesting property of hydrofluoric acid is its power of corroding glass, and it is often used in the arts to etch designs upon glass, which is done in the following manner.

Experiment 2.—Thoroughly warm a small square of window glass, and then rub over one side of it with a piece of wax, so as to make a smooth coating. When the glass has become cold and the wax hard, scratch any design with a pointed instrument on the coated side of the plate, taking care to cut quite through the wax. The glass plate thus prepared is then laid with the waxed side downwards, so as to cover the leaden basin in which the hydrofluoric acid is being prepared as directed in Experiment 1.

If, after being exposed to the acid fumes for a few minutes, the plate is removed and cleaned from the wax with a rag and a little turpentine, it will be found corroded wherever the marks have been made, and the design will remain engraved on the glass.

CHAPTER II.

NON-METALLIC DIADS.

(Oxygen, Sulphur, Selenium, Tellurium.)

OXYGEN.

Symbol, O = 16. Formula, O₂.

OXYGEN is the most abundant of the elements, for it is probable that not less than two-thirds of all that portion of our globe of which we know anything consists of it. About one two-billionth part of this oxygen exists in the atmosphere in a free, or uncombined state; the remainder occurs in combination with various other elements. It constitutes eight-ninths of the weight of water, and about one-half of the weight of the chief rocks, and it is an important constituent of all animals and vegetables.

In the pure state it is a gas of specific gravity 16, that is, about one-tenth heavier than common air. By extreme cold it is liquefied. It is colourless, tasteless, and inodorous, and is very slightly soluble in water, one hundred volumes of water dissolving about three volumes of the gas at a temperature of 15° C. Its compounds are called *oxides*, and oxides of all the elements, except fluorine, are known. The gas is magnetic; for a balloon filled with it is attracted by the poles of a powerful magnet.

Oxygen cannot readily be prepared from air, and in the great majority of its most abundant compounds it is held in combination by a force too strong to be easily overcome; but a certain number of oxygen-compounds are known, which are so unstable, that either by heat, or by some other agency, the element may be displaced and collected in a pure state.

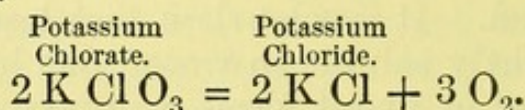
PREPARATION.

Oxygen was first prepared by Dr. Priestley, in 1774, from mercuric oxide, by the process already described (page 62). It is also obtained when water is decomposed by electricity (page 12).

Experiment 1.—Into the same apparatus that was used for heating mercuric oxide, introduce 100 grains of *potassium chlorate*, and heat it carefully with the spirit-lamp, or gas-lamp. The salt will soon melt, and afterwards boil. As soon as the boiling commences the flame must be lowered, to prevent the mass from frothing over. When the liquid thickens, if some of the substance should be found adherent to the colder parts of the tube, approach it with the flame of the lamp until it is again melted down. The gas may be collected over the pneumatic trough, just like hydrogen; but when it ceases to come off after strong heating, the delivery tube must *immediately* be removed from the water.

Instead of using pure potassium chlorate, it is much better to mix the salt with about one-tenth of its weight of *carefully-dried* manganese peroxide. The oxygen is then given off at a lower temperature and more easily, but the manganese peroxide remains unaltered after the experiment.

The following equation shows the change which takes place in this important reaction :



Potassium chlorate contains for every one hundred grains nearly forty grains of oxygen chemically combined; by the application of heat, these become free and escape. Mercuric oxide contains nearly eight per cent. of oxygen; therefore the former will yield five times more oxygen than the latter. If phials of twelve ounces' capacity are selected for receiving the gas, we shall be able to fill five of them, and shall have in each about eight grains, or nearly twenty cubic inches, of oxygen.

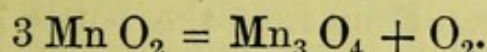
Potassium chlorate may, under some circumstances, particularly when treated with sulphuric acid, occasion

very dangerous explosions; but no danger is to be apprehended from the application of it in the manner above directed.

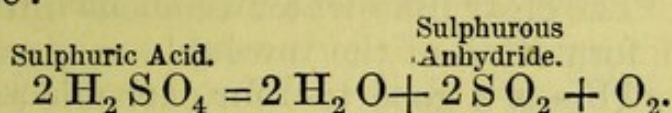
Experiment 2.—Add warm water to the salt remaining in the test-tube after the expulsion of the oxygen, and place the tube in a warm place until the salt is dissolved; evaporate the solution gradually, when small cubic crystals of potassium chloride will be deposited. The potassium chlorate crystallizes in thin tables or plates, the heated mass in cubes; this difference in the form of the crystals indicates that, by the heating of the former, an entirely new salt is formed. It is indeed one which no longer contains oxygen.

Oxygen may also be prepared by the following processes, which, however, are less convenient on the small scale than the foregoing:

1. By heating manganese peroxide to redness in an iron tube closed at one end:



2. By passing the vapour of sulphuric acid through a red-hot tube:



The sulphurous anhydride may be absorbed by passing the gas through lime. This, on the large scale, is a cheap process for preparing oxygen.

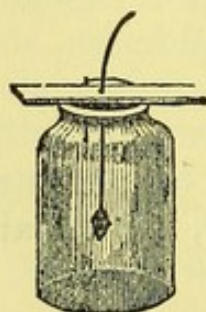
PROPERTIES.

Experiment 3.—Introduce a glowing shaving into a bottle of oxygen; it will kindle and burn for some time with great brilliancy and with a very dazzling flame, and then be extinguished. The same takes place when a piece of lighted tinder is fastened to a wire and suspended in the oxygen; the tinder burns with a lively flame, while, as is well known, it merely smoulders away in the open air. Oxygen, at a high temperature, combines eagerly with the component parts of wood and tinder, and heat and light are developed. When the combination is ended, and the oxygen is consumed, the combustion ceases. The product of the combustion, that is, the combination of the

wood with the oxygen, is also aeriform; but burning substances are extinguished in the newly-formed gas. If the bottle be rapidly whirled round, the gas formed by the combustion will escape, and atmospheric air will supply its place. Air contains free oxygen; and a kindled shaving will burn in it for some time, but far more slowly and less briskly than in pure oxygen; because common air contains only one-fifth part of oxygen. Accordingly, combustion proceeds much more rapidly and violently in oxygen than in atmospheric air.

Experiment 4.—Fasten a piece of charcoal to a wire, and kindle it in the flame of a lamp, and then introduce it into a bottle of oxygen; it will burn very vividly, and, if the charcoal had much bark on it, with beautiful sparks. If a piece of moistened blue litmus-paper be introduced into the bottle, after the combustion, it will be reddened; consequently an acid oxide has been formed from the charcoal and the oxygen; it is called *carbonic anhydride* or *dicarbon oxide*. Pour into the bottle a little *lime-water*

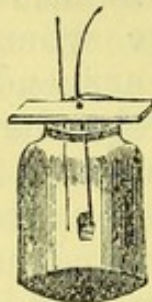
Fig. 43.



(which is prepared by shaking slaked lime with water for a short time and filtering). The clear lime-water becomes milky from the formation of the insoluble calcium carbonate. This is a good test for carbonic anhydride.

Experiment 5.—If some pieces of sulphur are fastened to a longer wire, kindled and suspended in a second bottle, they will burn with a beautiful blue flame. The gas formed from this union of sulphur and oxygen has a very irritating odour; when dissolved in water, it likewise turns litmus-paper red, and consequently it is of an acid nature. It is called *sulphurous anhydride*.

Fig. 44.



Experiment 6.—Take a small piece of phosphorus, which, on account of its inflammability, must be cut off under water from the stick, and place it, after it has been *well dried* between blotting-paper, in a scooped-out piece of chalk. Fasten the latter to a wire, and introduce it into a third bottle of oxygen.

Affix the wire to a cross-piece of wood, so that the chalk may hang a little below the centre of the bottle. If the

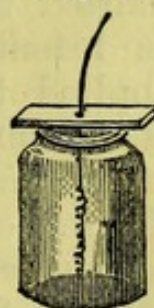
phosphorus be now touched with a hot wire, it will kindle and burn with a dazzling brilliancy, filling the bottle with a thick white smoke. This smoke consists of a chemical compound of oxygen and phosphorus; it reddens the blue test-paper, and consequently is also an *acid oxide*; it is called *phosphoric anhydride*. If the bottle be allowed to stand for a time, the smoke will sink to the bottom and dissolve in the water which remains there, which thus acquires an acid taste.

The two last experiments, and many others, can be more conveniently performed in *deflagrating spoons*, which are merely little ladles of iron or brass fastened at right angles on the end of iron wires, which pass through a cork and a piece of tin plate. The tin plate rests on the mouth of the bottle.

Experiment 7.—Heat a small piece of sodium in a deflagrating spoon till it begins to burn, and then plunge it in a bottle of oxygen. It will burn with great brilliancy, and with a yellow flame. A white oxide is formed, which, when dissolved in water, turns red litmus-paper to blue. Oxides of this kind are called *basic oxides* (page 100). Potassium yields a similar compound.

Experiment 8.—A piece of fine iron wire is so wound round a slate- or common lead-pencil, that, on the withdrawal of the latter, the wire may have a spiral form. Fasten the upper part of this wire, as in Experiment 6, to a cross-piece of wood, and place on the lower end of it a bit of tinder dipped in sulphur. When this is kindled, introduce the wire into the oxygen; the burning tinder heats the iron to redness, which then burns brilliantly, throwing out sparks. The iron, when red-hot, combines with the oxygen. The burnt or oxidized iron (iron scales) melts, and falls to the bottom in black globules, which are so hot that they are apt to melt into the glass, though it be partly filled with water. This black oxide of iron, Fe_3O_4 , is the same that is formed when steam is passed over red-hot iron in a tube. The inside of the bottle becomes covered during the experiment with a brick-red oxide,

Fig. 45.



Fe_2O_3 , which is called *ferric oxide*. Common iron-rust is ferric oxide. Ferric oxide is a basic oxide, although, being insoluble in water, it cannot be tested with litmus-paper. The black oxide is a *neutral* oxide. It is neither acid, nor basic.

These experiments show that there are three kinds of oxides :

1. Acid oxides, or anhydrides, which by combination with water yield acids.

2. Basic oxides, which by combination with water yield bases.

3. Neutral oxides.

The connection that these oxides have with the acids and bases has already been described (page 98).

Combustion.—The process which we call combustion, or burning, is nothing but rapid combination with oxygen, with evolution of light and intense heat. It is in fact only a particular case of oxidation. In all ordinary cases oxidation is attended with the production of heat, which we have before seen is a frequent result of the exertion and consequent loss of chemical force. But the oxidation of a certain weight of substance may take place rapidly or slowly, and the *intensity* of the heat will, of course, vary in like proportion, although the total *quantity* remains the same. The term combustion is generally limited to those cases of oxidation where light as well as intense heat is evolved. But the term may also be applied to similar combinations between other elements, as, for instance, between chlorine and hydrogen or metals.

Experiment 9.—Invert a tumbler over a small piece of lighted candle standing on a plate. In a short time the candle will go out, having removed a portion of the oxygen from the air of the tumbler. The inside of the tumbler will be seen to be covered with *moisture*, and if it is rapidly removed, and a little lime-water shaken in it, the latter will become milky, showing the formation of carbonic anhydride.

All ordinary kinds of fuel—coal, wood, gas, oil, &c.—contain carbon and hydrogen. This experiment shows the important fact that, *when they burn, the carbon becomes carbonic anhydride, and the hydrogen water*.

The phenomena of slow oxidation are very interesting and important.

Experiment 10.—Throw some iron filings, which have been moistened with water, into a bottle, so that they may adhere to its bottom. Invert the bottle with its mouth in a basin of water, and leave it in a warm place for a few days. The water will gradually ascend in the bottle, for the air becomes diminished in bulk to the extent of one-fifth; in fact, the whole of the oxygen is slowly absorbed by the iron and combines with it, forming *ferric oxide*, Fe_2O_3 . The gas remaining in the bottle is *nitrogen*, as is shown by plunging a lighted taper into the gas; the taper will be extinguished (p. 155). The *rusting* of iron is here seen to be its slow oxidation.

Experiment 11.—Bruise a little hay with water, stuff it in the bottom of another small phial, and proceed as in the former experiment. The same effect will be produced as with the iron filings; that is to say, the oxygen will be absorbed by the hay, and nitrogen will remain, mixed in this case with a little carbonic anhydride from the oxidation of the hay. The hay undergoes the process of *decay*, which is thus shown to be a slow oxidation. In both of these cases, heat is evolved, though it possesses but slight intensity. The heat of a *hot-bed* is due to the slow oxidation of the manure.

Respiration.—*Experiment 12.*—The heat of the animal body is maintained by a process of slow oxidation, analogous to that of the hay. Blow through a glass tube into a tumbler half full of lime-water. In a few minutes it will become turbid, showing that *carbonic anhydride* is contained in the expired air. Breathe into a cold and dry tumbler. It will become covered with *moisture*. Animals consume the oxygen of the air, and return carbonic anhydride and water to it. Fishes also require oxygen for respiration. They cannot decompose water, but they abstract the oxygen which it constantly absorbs from the air, returning carbonic anhydride in exchange. In water from which the oxygen has been expelled by boiling, fishes soon die of asphyxia.

Substances are more easily burnt when they are in a finely divided state than if they are solid; shavings are

easier to burn than billets of wood, and iron is easily combustible when in the form of filings.

Experiment 13.—Hold a small bar of iron in the flame of a spirit lamp; it will not burn; but file it and let the filings fall into the flame; they will be immediately ignited. Iron filings when sprinkled in a flame burn brilliantly.

Deflagration.—The term *deflagration* is applied to those cases of combustion where the body is burnt, not by the agency of *free* oxygen as that of the air, but by the aid of *combined* oxygen as it exists in substances which are easily decomposed, such as potassium chlorate, KClO_3 , and potassium nitrate, KNO_3 . Deflagration may be exemplified as follows:

Experiment 14.—Heat 20 or 30 grains of nitre in a test-tube with a spirit lamp, and when it is melted throw in some powdered charcoal or a fragment of sulphur; it will burn violently, being oxidized at the expense of the nitre.

OZONE, O_3 .

When certain kinds of electrical discharge occur in an atmosphere of pure oxygen, or even of air, a contraction of volume takes place, and some of the oxygen is converted into the curious gas called ozone (Gr., ὄζω, to smell), on account of its peculiar odour. This odour can be perceived in the neighbourhood of an electrical machine during its action. The specific gravity of ozone is found to be 24 ($\text{H}=1$), that of oxygen being 16; so that ozone is half as heavy again as oxygen. This fact is explained by the more than probable theory that every molecule of ozone contains three atoms of oxygen instead of two, as in ordinary oxygen gas.

Ozone has been obtained in the pure state as a blue liquid which yields a sky-blue vapour, but it is generally met with diluted greatly with oxygen or air. It is an amazingly powerful oxidizing, bleaching, and disinfecting agent, a property which is due to the ease with which it gives out its third atom of oxygen. At a temperature of about 250°C . (482°F .), it is reconverted into common oxygen, $2\text{O}_3 = 3\text{O}_2$. It can be formed by several methods besides the one above mentioned.

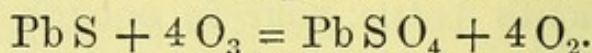
Experiment 1.—Take a stick of phosphorus, about an inch long, and scrape it with a knife, under water, so as to expose a new surface; then lay it in the bottom of a large wide-mouth bottle, pour in enough water to half cover it, and having loosely stoppered the bottle, set it aside in a warm place for about half an hour. White fumes of phosphorous acid will arise from the phosphorus, but will soon be absorbed by the water, and, simultaneously, ozone will be formed. The phosphorus combines gradually with part of the oxygen of the air in the bottle, while some of the remainder is converted into ozone, 3O_2 , becoming 2O_3 . Its presence may be shown by the use of what is called the “ozone test,” which is made by dissolving a few grains of potassium iodide in a tablespoonful of starch jelly, such as is prepared for use in the laundry, and smearing the latter on strips of paper. Introduce one of these test-papers in the bottle of ozone, and it will immediately become dark blue or violet, the colour being more or less intense according to the quantity of ozone. The iodine is liberated from the potassium iodide by the ozone, and it combines with the starch to form iodide of starch (*Experiment 5*, page 121).

Experiment 2.—Lay a piece of the test-paper in the bottom of a tumbler and gradually invert a bottle of ozone, prepared as in *Experiment 1*, as though in the act of pouring from the one vessel to the other. The ozone will stream out and occupy the tumbler, and come in contact with the test-paper, which will be immediately affected. This experiment shows that ozone is heavier than air.

Experiment 3.—Suspend a moist silver coin in a bottle of ozone. In a few minutes it will be covered with a grey deposit of silver oxide, thus showing the remarkable oxidizing power of ozone.

Experiment 4.—Place a few fragments of sulphide of iron in a glass or jar, and pour upon them a little water and a few drops of sulphuric acid; hydrogen sulphide, a gas of an extremely disagreeable odour, will be evolved. Moisten a strip of paper with a solution of lead acetate, and bring it in contact with the escaping sulphuretted hydrogen, and the paper will become dark brown or black, from the formation of *lead sulphide*, PbS . If the

paper be then plunged into a bottle containing ozone it will again become colourless. The ozone oxidizes lead *sulphide*, which is black, into lead *sulphate*, which is white :



Experiment 5.—Moisten a glass rod with a strong solution of ammonia, and introduce it into a bottle of ozone. White fumes will be abundantly formed, which consist of *ammonium nitrite*, NH_4NO_2 .

Experiment 6.—Moistened litmus-paper is immediately *bleached* when introduced into a bottle of ozonized air. The bleaching and disinfecting of bodies by ozone are owing to their oxidation. Ozone is very commonly found in the atmosphere, formed perhaps by the action of electricity. It cannot be doubted that this atmospheric ozone has important functions to fulfil in the economy of nature.

Experiment 7.—Strips of paper moistened with ozone test may be exposed to the air for a few hours. They will frequently be found to have turned blue, especially in country places.

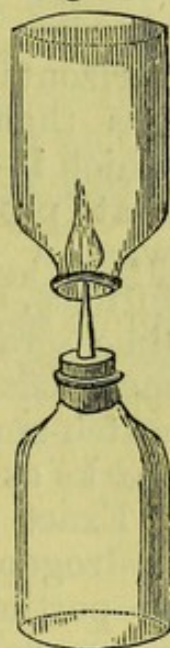
WATER, H_2O .

Water does not exist in nature in a perfectly pure condition. Even fresh water contains small quantities of solid impurities, of which the carbonates and sulphates of calcium and magnesium are among the most important and general. Natural water also holds in solution small quantities of oxygen, nitrogen, and carbonic acid ; but besides these general impurities, there are others which occur only in particular instances, such as iron, sulphuretted hydrogen, iodine, &c., which are present in certain springs, and which give to them a medicinal value, as is the case with the waters of Harrogate, of Cheltenham, and of Tunbridge Wells. Rain-water, collected in a vessel placed in a large open space, such as a field, is the purest kind of *natural* water that can be obtained, and may be used by the student, for most purposes, in lieu of *distilled* water, in cases where the latter cannot easily be obtained.

The composition of water was first determined by Cavendish ; he proved that it is the sole product of the combustion of hydrogen.

Experiment 1.—That water is really formed by the burning of hydrogen can easily be shown by inverting a clear glass bottle over the hydrogen flame; the glass soon becomes clouded over, because the water, which in consequence of the heat is generated in the form of steam, condenses in small globules on the cold sides of the bottle.

Fig. 46.



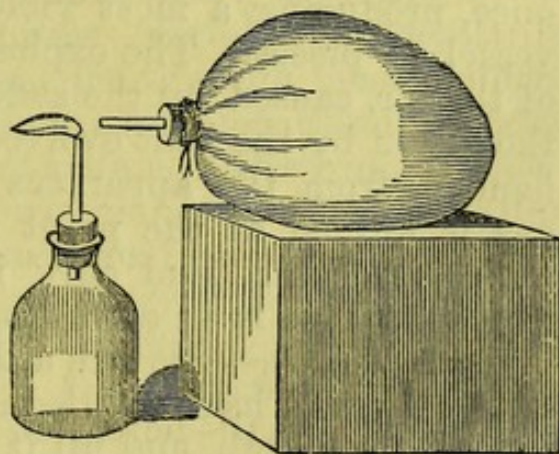
Experiment 2.—The extraordinary degree of heat developed by the chemical union of oxygen and hydrogen may be shown by the following experiments. Insert into the opening of a large pig's bladder, which has been softened by soaking in water, the broken-off neck of a flask, and bind it firmly round with a string. Then select two perforated corks, fitting this neck. One cork is connected with a bent glass tube, conducting the oxygen from the apparatus in which it is evolved into the bladder, which soon becomes filled with it. When this operation is finished, replace the first cork by the second, having a glass tube adapted to it only a few inches long, and drawn

Fig. 47

out to a point at its outer end, and provided with a wax plug pressed upon the opening. A glass tube may be formed into a jet by heating it in the flame of a spirit lamp, constantly turning it round at the same time, till it becomes so soft at the desired place as to be easily drawn out. Break it at the slender part, and hold it in the flame for some moments, until the sharp edges are rounded off by incipient melting. It would be more convenient, though somewhat more expensive, to substitute for the above contrivance a small brass stop-cock, provided with a jet.



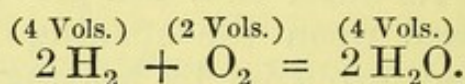
Fig. 48.



The bladder thus arranged and filled with oxygen is now placed upon a block, at such a height that the point of the glass tube

shall be on a level with the hydrogen flame, produced as explained in a former experiment. Press upon the bladder with the hand, and the oxygen will escape, blowing into the hydrogen flame, which then takes a horizontal direction. This flame has but little brilliancy, less than the hydrogen flame alone, notwithstanding which it affords the greatest heat yet known, except the heat produced by electricity. Hold in it a platinum wire, a metal which cannot be melted in the hottest furnace, and it will melt like wax; hold in it a piece of lime scraped to a fine point, and it will emit light of the most dazzling splendour. This is called the *lime light*. A watch-spring or a fine iron wire burns in it, throwing out sparks as in oxygen.

Exact experiments have shown that two measures of hydrogen unite with one measure of oxygen, the same proportion as that in which the gases are obtained when water is decomposed by electricity (page 12). The result of the combination is steam, which condenses into water. But two measures of hydrogen and one of oxygen do not yield three measures of steam; they afford two measures only :



Thus the two gases condense one-third by chemical union.

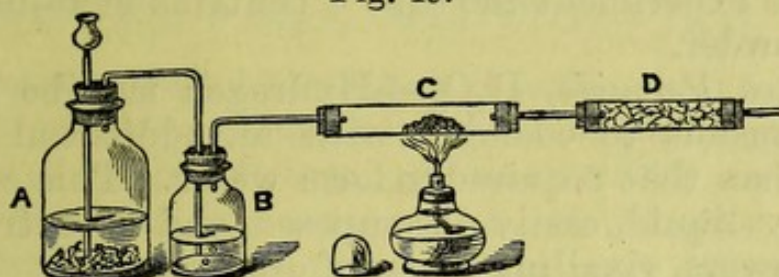
If the hydrogen and oxygen were mixed together and then ignited, the whole mass would combine together at once, producing a most violent report, and bursting the vessel to pieces. The explosion is due to the concussion of the air, caused by the sudden and violent expansion of the steam by the heat evolved. No danger is to be apprehended from the apparatus described, as the explosive gas is formed at the point where the oxygen meets the hydrogen flame, and only in small quantities at once. The apparatus is an *oxyhydrogen blowpipe* on a small scale.

Experiment 3.—Procure a soda-water bottle fitted with a good cork, and having filled it with water, invert it in the pneumatic trough, and fill it two-thirds full of hydrogen, and one-third with oxygen; cork it securely, and allow it to remain with its mouth immersed in the water

for five minutes, or shake it so that the gases shall inter-mix. If the bottle be then uncorked, and a light quickly applied to the mouth, the gases will combine with a loud explosion. Owing to the small quantity of the gaseous mixture employed there is little danger in this experiment; but to prevent a chance of accident it is advisable, before applying the light, to fold the bottle in a cloth, so as to secure the fragments of glass, should the bottle be broken by the explosion.

The composition of water by volume is thus easily illustrated; but to show directly the relative *weight* of its constituents, the following rather complicated arrangement may be used:

Fig. 49.



A is a hydrogen apparatus, and it is connected by a bent tube with the bottle B, containing sulphuric acid. B is connected in turn to the two wide tubes C and D. C contains a small heap of black copper oxide, CuO , heated with a spirit or gas lamp, and D is filled with fragments of calcium chloride.

Some granulated zinc having been placed in the bottle A, a mixture of sulphuric acid and water is poured down the long funnel called a thistle funnel. Hydrogen is evolved and passes by the bent tube through the sulphuric acid in the bottle B, and is thus dried, for sulphuric acid eagerly combines with and retains any watery vapour that may be carried over by the hydrogen. The dry gas, after bubbling through the sulphuric acid, passes through the wide tube C, and comes in contact with the copper oxide. When the common air is expelled, the tube C is heated. The hydrogen then combines with the oxygen of the copper oxide and forms water: $\text{CuO} + \text{H}_2 = \text{Cu} + \text{H}_2\text{O}$. The latter, as fast as it is formed, is driven forward and absorbed by the calcium chloride—a substance already mentioned as a

powerful absorbent of water—in the tube D. Now if the tube D be weighed before and after the experiment, any additional weight it may have gained will of course be owing to the water formed during the experiment, and thus its amount may be readily determined; but if the tube C containing the oxide of copper be also weighed, any *decrease* in weight will represent the quantity of oxygen which has been required to form the water obtained. In this way it may be determined that for every *nine* grains of water produced, the oxide of copper loses *eight* grains of oxygen: the difference is of course due to the hydrogen, and represents the relative weight with which it combines with oxygen to form water. At the end of the experiment the tube C contains metallic copper, in fine powder.

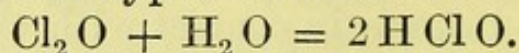
Hydrogen Peroxide, H_2O_2 .—Hydrogen may be made by indirect means to combine with an additional atom of oxygen than that required to form water. This compound is an oily liquid, easily decomposed, and of extreme oxidizing powers, rivalling those of ozone.

COMPOUNDS OF OXYGEN AND CHLORINE.

Oxides of Chlorine.	Corresponding Acids.
Cl_2O Chlorine Monoxide, or Hypochlorous Anhydride.	Hypochlorous Acid, $HClO$.
Cl_2O_3 Chlorine Trioxide, or Chlorous Anhydride.	Chlorous Acid, $HClO_2$.
ClO_2 Chlorine Tetroxide, or Chlorine Peroxide.	

It is not necessary for the beginner to study all the above compounds. Many of them are highly dangerous substances, and all are difficult to prepare. Chlorine peroxide is not an acid oxide.

Hypochlorous Anhydride, Cl_2O , is a colourless gas easily condensable to a red explosive liquid by cold. Combined with water it yields hypochlorous acid:

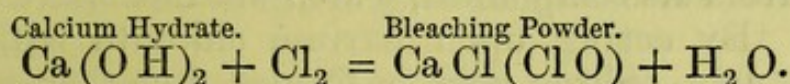


This acid corresponds to the series of salts called *hypochlorites*:

$HClO$, Hypochlorous Acid.
 $NaClO$, Sodium Hypochlorite.
 $Ca(ClO)_2$, Calcium Hypochlorite.

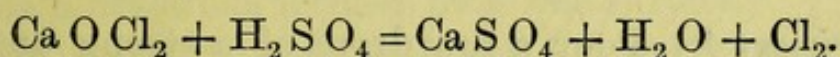
The substance commonly called chloride of lime, or bleaching powder, is intermediate in composition between the chloride and hypochlorite of calcium (CaCl_2 and $\text{Ca}(\text{ClO})_2$). Its composition may be represented by the formula CaOCl_2 or $\text{Ca Cl}(\text{ClO})$, or as a compound of calcium chloride and hypochlorite,— $\text{Ca Cl}_2 + \text{Ca}(\text{Cl O})_2$. Bleaching powder is prepared by acting on slaked lime (calcium hydrate) with chlorine.

Experiment 1.—Throw a little slaked lime into a bottle of chlorine, stoppered loosely. In a short time the chlorine will disappear, the lime will absorb it, and be converted into bleaching powder :



The manufacture of bleaching powder on a large scale is carried on by passing chlorine through chambers in which layers of slaked lime are spread. The process is complete when the lime ceases to absorb chlorine.

Experiment 2.—Pour a few drops of very dilute hydrochloric acid on a little bleaching powder. Effervescence will be produced by the escape of gaseous hypochlorous acid, which has an odour somewhat resembling that of chlorine. A piece of moistened litmus-paper, if it be immersed in the jar, will be bleached. If sulphuric acid is added *in excess*, chlorine is set free :



This is a simple and effectual way of preparing chlorine for disinfecting purposes.

In the ordinary process of bleaching, the materials to be bleached are dipped into a solution of bleaching powder, and then passed through a dilute acid ("soured"), which liberates the hypochlorous acid. The acid is an even more powerful bleaching agent than chlorine.

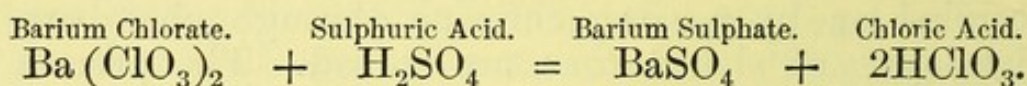
The peculiar odour which bleaching powder apparently possesses is due to hypochlorous acid, which is slowly and continuously formed from it by the carbonic acid of the air, so easily is the compound decomposed.

Chlorine peroxide, ClO_2 or Cl_2O_4 , is a dark yellow, easily decomposed gas, formed by the action of sulphuric acid on

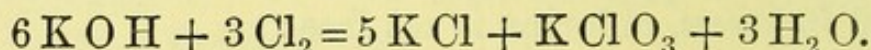
potassium chlorate. It may be prepared, and its explosive character safely observed, in the following way:

Experiment 3.—To two or three grains (not more) of potassium chlorate in a test-tube add a few drops of strong sulphuric acid, and gently warm. A peculiar crackling noise will be heard, while a coloured gas is slowly evolved. This gas is chlorine peroxide. If the tube be more strongly heated a loud explosion will often be produced from the sudden decomposition of the gas.

Chloric acid, HClO_3 , an oily liquid, easily decomposed, and of little utility. Its salts, especially the potassium salt, are of great importance. Chloric acid can only be prepared from its compounds, which are themselves readily obtained. By acting upon barium chlorate with dilute sulphuric acid, chloric acid is formed, together with the insoluble substance, barium sulphate:



Potassium chlorate is formed by a process analogous to that used for the preparation of bleaching powder. By passing chlorine through a cold dilute solution of potassium hydrate, a mixture of potassium chloride and hypochlorite is obtained; but if a hot solution be used, chlorate instead of hypochlorite is formed:



Experiment 4.—Pour a few drops of a strong and warm solution of potassium hydrate into a bottle of chlorine loosely stoppered. The chlorine will be absorbed, just as it was by calcium hydrate, while a white crystalline powder is abundantly formed which mainly consists of potassium chlorate, the chloride being very soluble. The chlorates are unstable bodies, and readily yield up their oxygen when heated; they therefore furnish convenient sources of that gas. It will be remembered that potassium chlorate is the salt commonly employed for the preparation of oxygen.

Experiment 5.—Powder a few grains of potassium chlorate; mix it gently with an equal quantity of powdered sugar, and touch the mixture with a rod dipped

in strong sulphuric acid. A brilliant combustion will ensue, owing to the ClO_2 set free from the chlorate.

Experiment 6.—If one grain of potassium chlorate is triturated in a mortar with one grain of sulphur, a sharp detonation will follow, the sulphur being oxidized at the expense of the salt.

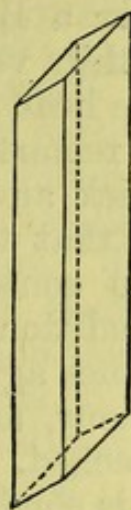
SULPHUR.

Symbol, S = 32. Formula of Vapour, S_2 and S_6 .

The well-known element sulphur, which, on account of its easy combustibility, is employed in the manufacture of matches, &c., has neither taste nor smell. It is found in nature in considerable quantity, mostly in a state of combination.

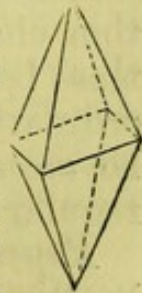
Experiment 1.—Sulphur is *fusible*. Heat two ounces of sulphur in small pieces in a small stone-ware crucible or Florence flask, over a lamp at a gentle heat; it is converted, at a temperature a little above that of boiling water, into a thin yellow fluid. If you pour some of it into cold water, you obtain again solid sulphur. If this, after being previously dried, is returned to the crucible, it will sink in the fluid mass, showing that solid is heavier than melted sulphur.

Fig. 50.



Experiment 2.—Sulphur may be crystallized. Let the crucible containing the melted sulphur stand till a crust has formed over the surface; break this quickly, and pour out the portion remaining fluid. Upon afterwards breaking the crucible, the cavity of the sulphur will be found lined with fine transparent crystals, in the form of lengthened pillars (Fig. 50), which are called monoclinic, or oblique rhombic prisms.

In different parts of the world, particularly in volcanic countries, such as Sicily and Iceland, large beds of sulphur (native sulphur) are not unfrequently found, and, in these beds, fissures and cavities studded with the most beautiful crystals. These native crystals have a very different form from those prepared by fusion. They appear like two pointed four-sided pyramids, applied base to base (Fig. 51);



such a form is called an acute octahedron, because contained under eight acute triangles. Thus sulphur, like carbon in the diamond and graphite, assumes two different crystalline forms; it is *dimorphous*.

Experiment 3.—Although sulphur is insoluble in *water*, many liquids are known which dissolve it; the one most convenient for this purpose is carbon disulphide. Pour a little of this liquid into a small clean bottle, and add a few pieces of sulphur. Close the bottle, and shake it at intervals for an hour or two. Then filter the liquid from the undissolved sulphur into a small basin, and allow it to evaporate; as it does so *crystals* of sulphur of the form shown in Fig. 51 will appear. Their shape differs from those obtained by fusion, but is similar to the native crystals.

Experiment 4.—Put about two ounces of fragments of roll-sulphur into a dry Florence flask, and place the latter on a retort stand. Heat it gently at first, but afterwards over a Bunsen's gas flame. The sulphur will first melt to a thin yellowish liquid. When it is all melted, and when the heat is increased, the character of the liquefied sulphur is remarkably altered; it loses its mobility, and becomes thick and pasty, and dark in colour like treacle; so much so that the flask may be quite inverted without danger of the contents running out. As the heat continues, this condition is likewise changed, and soon the sulphur becomes again a *thin* liquid, which retains, however, its dark colour, and thus it continues until the boiling-point is reached. If the sulphur be thrown into water while in this condition, a substance entirely different in appearance from ordinary sulphur is obtained in the form of a reddish-brown, soft, elastic mass, which serves admirably as a medium for taking impressions of seals, medals, &c. But this change is not a permanent one; in a short time the mass becomes hard and brittle, and is soon reconverted into ordinary sulphur. This plastic form of sulphur is said to be *amorphous*, a term applied to other bodies having no regular form, such as gum, pitch, glue, &c.

Experiment 5.—If, instead of pouring out the sulphur as in the last experiment, the heat be further applied, the sulphur will soon give out a rich dark-red vapour, which

gradually occupies the whole of the upper part of the flask: thus sulphur is *volatile*, and may, like water, assume all the three states of aggregation (solid, fluid, and aeriform). Solid sulphur has a specific gravity of about 2; that is, twice as heavy as water. Sulphur vapour has a specific gravity of 96 ($H = 1$), but at a very high temperature (1000°C.) it trebles its volume, and has a specific gravity of only 32. The formula for the ordinary vapour is, therefore, S_6 , and for the vapour at high temperatures, S_2 . Within the flask the vapour of the sulphur is transparent, and has a reddish-brown colour; but after escaping it appears as a yellowish smoke, being condensed by the cold air into a dust of solid sulphur. If this vapour be conducted into a glass jar, immersed in cold water, the sulphur condenses in it in the form of a soft yellow powder, known in commerce by the name of *flowers of sulphur*. The process by which a volatile substance is evaporated and condensed again into a solid is called *sublimation*. In distillation, the vapour is condensed into liquid (the distillate); in sublimation, into a solid (the sublimate).

The flowers of sulphur of commerce are obtained by vaporizing the impure native sulphur in iron or earthen retorts, and conducting the vapour into large chambers built of brick or stone, where it condenses as a crystalline powder, and lodges on the walls of the chambers. The interior of the chamber is sometimes allowed to become sufficiently hot to melt the deposited sulphur, which therefore runs down in the melted condition, and collects on the floor. While still liquid it is drawn off and solidified in wooden moulds, and so forms *roll-sulphur*, or *brimstone*.

Experiment 6.—Fill a test-tube one-third full of solution of caustic soda; add to it as much flowers of sulphur as can be taken up on the point of a knife, and boil the mixture for some time; a part of the sulphur will be dissolved, imparting to the liquid a yellowish-brown colour. The clear liquid is now decanted, diluted with water, and hydrochloric acid added to it; it will immediately assume a milky appearance, owing to the separation of the sulphur in the form of an exceedingly fine powder, which is so light that a considerable time must elapse before it will subside. Collect the powder on a filter, wash it with

water, and dry it at a gentle heat. It is called *milk of sulphur*, or *precipitated sulphur*, and is sulphur in its finest state of subdivision. Precipitated sulphur has a pale yellowish tint, but on being melted it becomes distinctly yellow, owing to the union of the individual particles into a larger mass.

Experiment 7.—If sulphur be heated in a vessel with free access of air, as in an iron spoon, or be touched by some red-hot body, it *burns* with a blue flame; that is, it unites with the oxygen of the air, under the phenomenon of fire, and forms with the oxygen, as has been previously shown, a suffocating gas, sulphurous *anhydride* (SO_2).

This property, which belongs to sulphur, of igniting at a very moderate heat, is the reason of its being so commonly used for *kindling purposes*. By means of it, other bodies more difficult of combustion may be heated to the temperature at which they can continue to burn (matches, gunpowder, fireworks, &c.). The kindling of a simple coal-fire well illustrates how, by gradual transition from easily inflammable materials to those of more difficult ignition, the latter are finally brought to that degree of heat at which they will ignite and continue to burn.

Sulphur is an energetic element, and has, like oxygen, a powerful affinity for other elements.

Experiment 8.—Boil some sulphur in a flask, and expose a very thin copper-plate to the brownish vapour; the copper will glow vividly for some moments, lose its red colour and flexibility, become grey and brittle, and weigh more than before. The newly-formed grey crystalline body is copper sulphide, CuS . Both elements have intimately combined, and in fixed proportions. The properties of the sulphur, as well as of the copper, have entirely disappeared. The great heat produced is a consequence of the chemical combination.

In a similar manner many other metals may be converted into sulphides. We find many, however, already formed in the earth, and miners call them glance, blende, or pyrites. The pyrites having the lustre of brass, and found in almost all coal, is sulphide of iron; red cinnabar is a sulphide of mercury, &c. The sulphide of copper, artificially prepared as above, occurs also as an ore.

Experiment 9.—Mix three-fourths of an ounce of iron filings, half an ounce of flowers of sulphur, and one-fourth of an ounce of water, in a small vessel, and put it in a warm place; the mass becomes heated, the water evaporates, and in half an hour a black powder will be obtained, in which no particles of iron or of sulphur will be perceived; a chemical compound, *ferrous sulphide*, is formed. If the two substances be mixed together without water, no combination will take place, unless they be heated to redness.

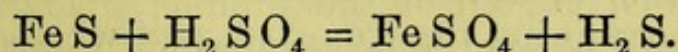
Experiment 10.—Melt an ounce of sulphur in a covered earthen crucible, and then gradually throw in rather more than half an ounce of iron filings, closing the crucible after each addition. The iron will glow and combine with the sulphur just as the copper did in Experiment 8. The compound formed is the same as that obtained in Experiment 9, namely, *ferrous sulphide*, FeS . When larger quantities are required it is better to buy than to make it.

HYDROGEN SULPHIDE, OR SULPHURETTED HYDROGEN, H_2S .

Experiment 1.—Put half an ounce of ferrous sulphide (FeS) and half an ounce of diluted sulphuric acid into a small bottle, and quickly stop the bottle with a cork, to which a bent glass tube is adapted. Introduce the longer limb of the tube into a bottle filled with cold water. The atmospheric air contained in the flask and tube first passes over, followed by a very offensive gas, which dissolves in the water, to which it likewise imparts its fetid odour of rotten eggs. This gas is called *hydrogen sulphide* or *sulphuretted hydrogen*. The following change takes place:



Fig. 52.



When the disengagement of the gas ceases, add some diluted sulphuric acid, that the gas may again be generated. The water is known to be saturated with the gas when, on shaking the bottle, the finger by which the opening is closed is no longer sucked in, or, more correctly speaking, pressed in. One measure of water contains about two and a half

measures of gas in a saturated solution. It is put up in small well-stoppered bottles. If the air be admitted, the solution becomes turbid, owing to the oxygen of the air uniting with the hydrogen of the sulphuretted hydrogen, forming water, and the consequent liberation of the sulphur as a fine powder.

If, during the evolution of the gas, the bottle of water be removed, the gas issuing from the tube can be ignited by a match; it burns with a blue flame, and its nauseous odour is no longer perceptible, but is replaced by the well-known odour of burning sulphur. Both constituents unite with the oxygen of the air, the sulphur forming sulphurous anhydride, and the hydrogen, water.

The inhalation of sulphuretted hydrogen is detrimental to health; hence care should be taken to avoid it. When experimenting with it, it is best to do so where there is a free circulation of air.

Hydrogen sulphide turns blue litmus-paper red; it also combines with bases, and hence it may be regarded as an acid.

Experiment 2.—Drop some hydrogen sulphide water upon a bright silver or copper coin, and upon a piece of lead and iron. The first three metals tarnish quickly, and finally become black; they combine with the sulphur, forming dark *metallic sulphides*; the iron, on the contrary, undergoes no change.

Experiment 3.—Put into one test-tube a small portion of litharge, into another some ignited iron-rust, and pour upon them a solution of this gas. The yellow litharge, oxide of lead, becomes immediately black, an exchange of elements takes place, the hydrogen sulphide gives its sulphur to the lead of the litharge, and receives in return the oxygen of the latter. Accordingly, lead sulphide and water are formed, and the offensive odour disappears. In the vessel containing the iron-rust, neither the colour nor the smell is changed—a proof that no chemical change has taken place.

Experiment 4.—Repeat the same experiment with a small crystal of sugar of lead instead of the litharge, and some green vitriol instead of the iron-rust, together with a few drops of acetic acid, these salts having been previously

dissolved in a large quantity of water; the result will be the same as in the former experiment. Sugar of lead is lead acetate; the salt of lead is converted into sulphide, which subsides sooner or later as a black precipitate. When this solution is extremely diluted it is only rendered brown. Acetic acid is set free, and remains in solution.

Most of the metallic sulphides are *insoluble* in water; hence hydrogen sulphide is peculiarly adapted for precipitating metals from their solution, so that they can be separated and collected by filtration. If hydrogen sulphide is passed through a solution of acetate or sulphate of copper, sulphide of copper will be precipitated, and can be separated by filtration from the acid. Some of the sulphides do not possess a black colour; sulphide of antimony has an orange-red, sulphide of arsenic a yellow, and sulphide of zinc a white colour. On this is partly based the application of hydrogen sulphide as a test; that is, as a means of detecting many metals. Wine containing lead is blackened by hydrogen sulphide.

Many metals are precipitated from their acidified solutions as sulphides by the addition only of sulphuretted hydrogen; for example, copper, silver, gold, lead, mercury, tin, antimony, and arsenic; and others are not precipitated until a base is added; for example, iron, zinc, manganese, cobalt, and nickel. Hydrogen sulphide may accordingly be used to separate one metal from another; it is, therefore, an important means of separation in analytical chemistry. Lead paper is used for the detection of hydrogen sulphide, by which it is coloured brown or black. It is made by passing strips of paper through a weak solution of sugar of lead in water.

Finally, it remains to be stated that this gas occurs also in some *mineral waters*, as may be recognised by the smell and taste. Many of these springs—for instance, the celebrated springs of Harrogate and Aix-la-Chapelle—are resorted to by invalids, and are called *sulphur springs*. A rotten wooden water-pipe may convert an otherwise potable water, if it should contain calcium sulphate, into a nauseous sulphuretted water; but by clearing out the well, and laying down new pipes, the water may be rendered completely odourless and good.

COMPOUNDS OF SULPHUR AND OXYGEN.

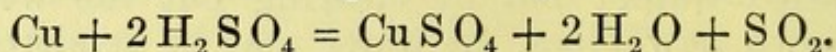
Sulphur combines with oxygen in two proportions, forming the oxides SO_2 and SO_3 ; they both unite with water to form acids. SO_2 is hence called sulphur dioxide, or sulphurous anhydride, and SO_3 sulphur trioxide, or sulphuric anhydride. The latter substance can only be prepared indirectly, whereas sulphurous anhydride is formed by the direct union of its constituents whenever sulphur is burnt in air or oxygen, as has been already mentioned.

Sulphurous Anhydride, SO_2 , and Sulphurous Acid, H_2SO_3 .—Sulphurous anhydride is a colourless gas of a pungent suffocating odour, well known as the smell of sulphur. One of its most characteristic properties is its power of bleaching animal and vegetable colours.

Experiment 1.—Suspend a red flower, such as a rose or peony, in an inverted glass jar, or other similar vessel, and introduce a piece of ignited sulphur. It will be found that the colour of the flower is slowly changed to white. Sulphurous anhydride, unlike chlorine, does not actually *destroy* the colouring matter, for by dipping the flower into dilute sulphuric acid the original colour is slowly restored. Straw, wool, silk, &c., are very commonly bleached with sulphurous anhydride, but the result of its action on these materials is far from permanent; for on exposing a bleached straw hat in the sunshine it again becomes yellow, or “sun-burnt,” as the effect is popularly termed; and on washing whitened flannel it resumes its original appearance.

Sulphurous anhydride can be formed in a variety of ways: one of the most convenient, when a quantity is required in a tolerable state of purity, is the following:

Experiment 2.—Select a thin-bottomed, well-made Florence flask, fit it with a delivery tube, and place in it half an ounce of thin copper, sheet or turnings, and two ounces of strong sulphuric acid. Apply a gentle but gradually increasing heat, and after a time sulphurous anhydride will be evolved, according to the equation:

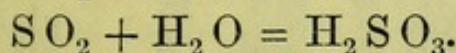


Sulphurous anhydride is more than twice as heavy as

air, so it may be collected in dry bottles by *displacement*; that is, by allowing the gas to pass to the bottom of a dry bottle. It will gradually displace the lighter air, just as water does when poured into a bottle, and as it is very soluble in water this mode is the most convenient.

Experiment 3.—Pour a little blue litmus solution into a bottle of sulphurous anhydride; it will be reddened, showing the acid character of the gas.

Experiment 4.—Dip the end of the delivery tube from which the gas is issuing into a little water; the gas is absorbed in large quantities and imparts to the water its characteristic taste and smell. One measure of water dissolves about forty measures of the gas. The solution obtained is one of sulphurous *acid*:

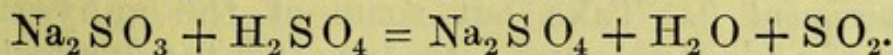


The salts corresponding to sulphurous acid are called *sulphites*:

$\text{H}_2 \text{S O}_3$, Sulphurous acid.
 $\text{Na}_2 \text{S O}_3$, Sodium sulphite.
 Ca S O_3 , Calcium sulphite.

Experiment 5.—When the gas ceases to be absorbed by the water, substitute for the latter a solution of sodium carbonate; this likewise absorbs the gas and forms with it sodium sulphite, while the carbonic acid is liberated with effervescence. If a sulphite be treated with a strong acid it is decomposed and sulphurous anhydride is evolved.

Experiment 6.—Place a few grains of sodium sulphite in a test-tube, pour upon it a little slightly diluted sulphuric acid, and gently warm. Sulphurous anhydride is given off and is readily identified by its odour:

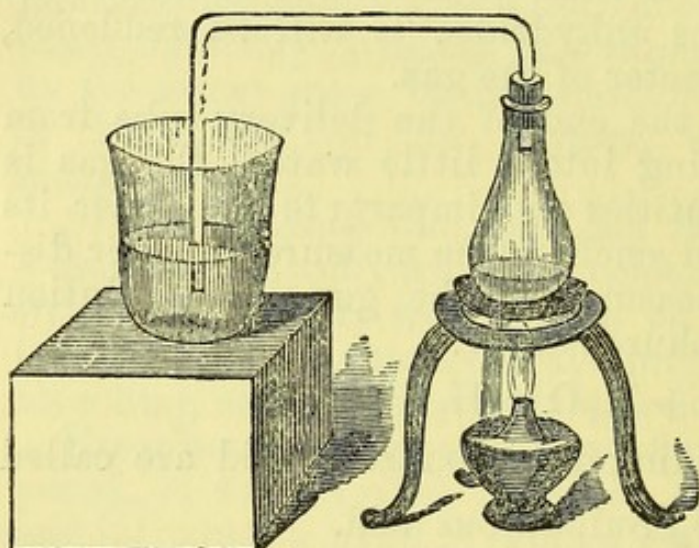


Experiment 7.—When the residue in the flask, in which the sulphurous anhydride was generated, has become cold, add water to it, and heat it gently to boiling until most of the residual mass is dissolved. The solution is dark and turbid, but after filtering it is of a beautiful blue colour, and transparent. If allowed to cool slowly, blue crystals of copper sulphate, $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ (blue vitriol), of considerable size will be formed. Silver and mercury comport themselves like copper and may be used instead.

Sulphuric Anhydride, SO_3 , and Sulphuric Acid, H_2SO_4 .

Sulphuric Anhydride.—*Experiment 8.*—Pour into a small flask, placed in a sand-bath over a tripod, half an ounce

Fig. 53.



of *fuming*, or Nordhausen sulphuric acid, and heat it gently till it boils moderately. Conduct the vapour through a tolerably wide glass tube into an empty flask, not shown in the cut, which is placed in a vessel filled with cold water. If the vapour be suffered to escape into the air, it appears in

thick white fumes, having a pungent acid smell; but if conducted into the flask, it is condensed into a liquid which solidifies into a glistening white solid mass. This is sulphuric anhydride. The distillation is stopped as soon as the boiling ceases, and the glass tube becomes too hot for the hand to bear. What remains in the flask no longer fumes; it has become common sulphuric acid. To cause this to boil, you must apply a much stronger heat than before, for it does not begin to boil till 640°F. (338°C.), while the anhydride boils at 115°F. (46°C.). This is the reason why the boiling ceases when the latter has escaped.

Experiment 9.—Take out some of the anhydride by means of a glass rod, and introduce it into a dry test-tube; it will fume violently, and after a time become fluid; *that is, it attracts water* from the air, and is thereby converted into Nordhausen sulphuric acid. On longer standing, it absorbs still more water, and ceases to fume; it thus becomes common sulphuric acid. By evaporation this water cannot again be removed.

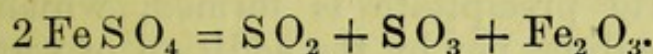
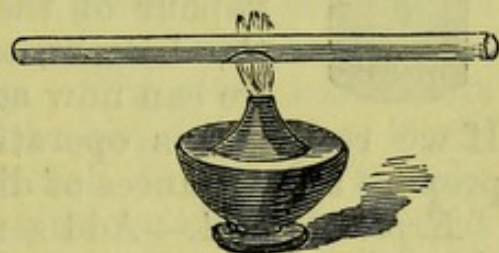
Experiment 10.—If anhydrous sulphuric acid be thrown into water, it is dissolved with a hissing noise and the *violent evolution of heat*.

Experiment 11.—It is likewise dissolved by common sulphuric acid, converting this into the fuming acid. Fuming sulphuric acid may therefore be regarded as a compound of the anhydride with ordinary sulphuric acid, $\text{H}_2\text{SO}_4, \text{SO}_3$.

Fuming, or Nordhausen sulphuric acid, is obtained by the distillation of green vitriol, or ferrous sulphate.

Experiment 12.—Put a crystal of green vitriol into a hard glass tube, and heat it; aqueous vapour escapes, and the green crystal becomes white (anhydrous). On further heating, the white colour passes into reddish-brown, and sulphurous and sulphuric anhydrides are evolved, while the iron remains as ferric oxide:

Fig. 54.



In preparing the fuming acid on a large scale, earthen retorts are used, and the anhydride is conducted into common sulphuric acid, which dissolves the sulphuric anhydride, and is thereby converted into the fuming acid, while the sulphurous anhydride escapes. It is a thick liquid, like oil, and is called Nordhausen sulphuric acid, because this city supplied Germany with it for centuries. It has the specific gravity of 1.9.

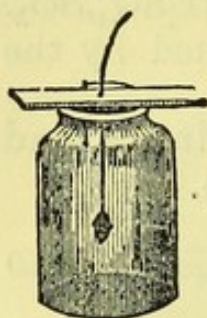
As long as the process of manufacturing sulphuric acid from green vitriol was the only one known, it was a very expensive article. A hundredweight is now obtained for the same sum that was formerly paid for two pounds.

Common Sulphuric Acid, H_2SO_4 .—Sulphur by burning is converted into sulphurous anhydride, SO_2 , and this combined with water gives sulphurous acid, H_2SO_3 . To convert this into sulphuric acid, H_2SO_4 , an additional atom of oxygen is necessary, and this additional oxygen is generally supplied at the expense of nitric acid.

Experiment 13.—Fasten some pieces of sulphur to an iron wire, inflame, and hold them in a capacious bottle containing a little water, until the blue sulphur flame is extinguished; the bottle becomes filled with a white smoke, which is recognised by its odour to be sulphurous acid. If you now introduce a shaving moistened with

nitric acid into the vessel, reddish-yellow fumes will immediately form around the wood, gradually filling the

Fig. 55.

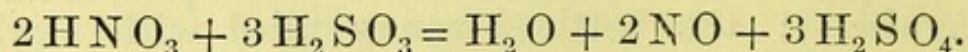


whole bottle. These fumes are nitric peroxide, and their evolution indicates that the sulphurous acid has withdrawn oxygen from the nitric acid, and has been oxidized and converted into sulphuric acid. After some time the bottle becomes clear again, because the vapour of the sulphuric acid formed sinks to the bottom, and dissolves in the water, and we can now again burn sulphur in the bottle.

If we repeat this operation several times, we can soon prepare a few ounces of dilute sulphuric acid.

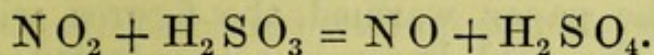
Experiment 14.—Add some drops of a solution of barium chloride to a portion of the acid liquid just obtained; a copious white precipitate is formed, which disappears neither by boiling, nor by the addition of water or nitric acid. This precipitate is *barium sulphate*, a salt quite insoluble in water and acids. Add one drop of the diluted acid to a wine-glassful of water, and add to this a little more of the solution of barium chloride; even at this great dilution a perceptible cloudiness will be produced. A solution of barium chloride, or nitrate, is the most certain test for detecting sulphuric acid and sulphates.

The manufacture of common sulphuric acid on a large scale is conducted on the same principle as in the last experiment but one. The sulphur is burnt on a sort of hearth, and the sulphurous anhydride thus formed is carried by the draught of air into an immense leaden chamber, the atmosphere of which is kept moist by jets of steam. The sulphurous anhydride thus becomes sulphurous acid. A little nitric acid (or sodium nitrate and sulphuric acid, which yield nitric acid) is placed in a basin, supported above the burning sulphur, and the heat drives it in vapour into the chamber. In the leaden chamber this nitric acid immediately oxidizes a portion of sulphurous acid into sulphuric acid, and is itself reduced to nitric oxide, NO:



Now we shall hereafter learn that nitric oxide is a

colourless gas which, on contact with air, takes up oxygen and forms the brown gas, nitric peroxide, NO_2 . $\text{NO} + \text{O} = \text{NO}_2$. This change immediately takes place in the leaden chamber, and as fast as the NO_2 is formed it is again reduced to NO by another portion of sulphurous acid :



Again the NO takes oxygen from the air of the chamber, and becomes NO_2 , and again a new portion of sulphurous acid is oxidized into sulphuric acid, so that the process, after the first stage, is continuous. The nitric oxide acts as a *carrier*, taking oxygen from the air, and giving it out again to the sulphurous acid, which is unable to take oxygen direct from the air (except very slowly), so that theoretically there is no limit to the quantity of sulphurous acid which may be oxidized by a given weight of nitric acid. As fast as the sulphuric acid forms, it condenses and settles down on the floor of the leaden chamber. From time to time it is drawn off and concentrated by evaporation, first in leaden, and afterwards in glass or platinum vessels. The acid thus obtained has almost exactly the composition of true sulphuric acid, H_2SO_4 .

The above account of the sulphuric acid manufacture must only be regarded as an approximation to the truth, for the process is in fact extremely complex. If the quantity of steam be deficient, white crystals of complex composition are apt to form. On the addition of water, however, they are decomposed, and sulphuric acid formed from them.

Experiment 15.—Let some sulphuric acid remain in an open flask exposed to the air; it will increase every day in weight, *for it very eagerly attracts water from the air*. After standing for some months in a damp place it will become two or three times heavier than before. Some substances, especially gases, are dried by means of sulphuric acid.

Experiment 16.—A piece of wood introduced into sulphuric acid becomes black, and is charred just as when it is exposed to the flame of a candle. The sulphuric acid seizes upon its hydrogen and oxygen, which combine to form water, and the carbon is left behind. Wood may be charred

in this way, in order to protect it from decay in moist situations. In the refining of lamp-oil, the mucilage of the oil is charred by sulphuric acid. *Sulphuric acid chars and destroys many vegetable and animal substances.* If figures are inscribed on paper with very dilute sulphuric acid, and the paper warmed, the figures will come out black. As the acid becomes concentrated, it acts upon and chars the paper.

Experiment 17.—Pour a drop of oil of vitriol upon paper; decomposition takes place slowly, but it will take place instantaneously if a drop of water is added, because *water and sulphuric acid unite together with the evolution of strong heat.* For this reason, when sulphuric acid comes in contact with the skin, it should first be wiped off with dry paper or cloth, and then be immediately washed with a great quantity of water.

Experiment 18.—Pulverise a small quantity of *indigo*, and form a thin paste of it with *fuming sulphuric acid*. After a few days add to it some water, and you obtain a deep blue liquid—solution of indigo. With this solution wool may be dyed of a fine blue colour (Saxon-blue). Common sulphuric acid dissolves indigo only imperfectly. Indigo, although a vegetable substance, is not carbonised by sulphuric acid, thus forming an exception to the general rule.

Experiment 19.—To half an ounce of copper oxide add two ounces of water, and then gradually two-thirds of an ounce of sulphuric acid, and put it in a warm place; you obtain a blue solution, from which afterwards blue crystals will be deposited. The edges of these crystals are usually obtuse, giving to the narrow sides a roof-like appearance. Copper oxide combines with the acid, forming copper sulphate (blue vitriol), a *soluble* salt. The following metallic sulphates are commonly called vitriols: ferrous sulphate (green vitriol); zinc sulphate (white vitriol); and copper sulphate (blue vitriol). Copper or iron vessels are cleaned more rapidly and made brighter by water to which some sulphuric acid has been added, than by simple water alone, because the oxide, which tarnished the vessels, is dissolved by the acid.

Experiment 20.—Put a small iron nail into a test-tube,

and drench it with twenty drops of common sulphuric acid ; it is not acted upon. But if you add a little water, about four or five times more than the acid, a brisk effervescence will ensue, and the iron will be dissolved. The strong acid may be heated to boiling in iron vessels without acting upon them, which is by no means the case with the diluted acid. In this experiment, hydrogen escapes, and iron sulphate remains in solution. Zinc behaves in a similar manner. Consequently, the *diluted acid* must be employed for dissolving such metals. But there are also metals which dissolve only in the stronger acid, with the aid of heat ; for instance, copper, silver, &c. This has been treated of under sulphurous acid.

SELENIUM, Se = 79, AND TELLURIUM, Te = 128.

These elements are very scarce and unimportant ; they occur combined with certain metals in some rare minerals. They both resemble sulphur in their properties, but partake more of the character of metals in appearance. Their compounds with hydrogen and oxygen correspond perfectly to those of sulphur.

CHAPTER III.

NON-METALLIC TRIADS.

(Nitrogen, Phosphorus, Boron.)

NITROGEN.

Symbol, N = 14. Formula, N_2 .

In the free state as gas nitrogen constitutes about four-fifths of the atmosphere, the remaining one-fifth being oxygen. In a state of combination it is an important constituent of organic structures. It is also found in the mineral kingdom, as in saltpetre or nitre, a salt occurring as an incrustation on the earth in India and elsewhere; from this salt the name *nitrogen* (generator of nitre) is derived.

The most convenient and simple method for the preparation of nitrogen consists in removing the oxygen from a confined portion of air by the ordinary process of combustion.

Experiment 1.—Fasten a piece of sponge to a wire, drop some alcohol upon it, and hold the wire in a vessel containing water so that the sponge may be some inches above the water. Then kindle the spirit, and immediately place an empty bottle over it, so that the mouth of it may dip into the water; the flame will soon cease burning, and some of the water will rise into the bottle, in proportion to the amount of air disappearing during the combustion. The consumed air was oxygen, which united

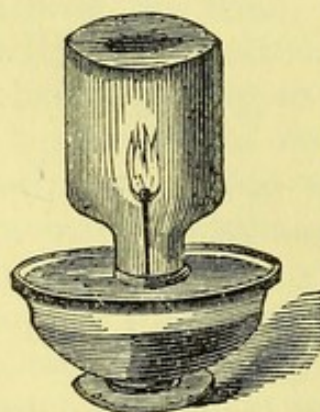


Fig. 56.

with the constituents of the alcohol. Close the bottle tightly with the finger, shake it briskly, and again open it below the water, when a little more water will enter.

The air which remains in the bottle is *nitrogen*; it is sometimes called *azote* (α , privative, and $\zeta\omega\eta$, *life*), from its inability to support respiration.

Other combustibles may be used for the purpose of abstracting the oxygen from air, such as sulphur, charcoal, &c., but phosphorus is the best; for the products of combustion of the other substances, being gaseous, are not easily got rid of, and consequently contaminate the nitrogen, whereas the phosphorus compound produced is solid, and moreover very soluble in water. Moreover phosphorus burns till all the oxygen is removed, which is not the case with many combustibles.

Experiment 2.—Place a small piece of dry phosphorus in a little basin floating in water. Ignite the phosphorus and place over it a bell-jar or the mouth of a wide-necked bottle. The phosphorus will burn until the whole of the oxygen is consumed. Dense white clouds of phosphoric anhydride (P_2O_5) are formed and fill the jar, but these after a time are precipitated in snowy flakes, and are rapidly dissolved by the water. The water will rise in the bottle, partly from the removal of oxygen and partly from some air having been expelled at the beginning of the experiment in consequence of the expansion produced by the heat. When the white fumes have entirely disappeared the bottle may be shaken so as to cause the basin to fill and sink, and the nitrogen transferred in the pneumatic trough to a series of smaller bottles.

Nitrogen is incombustible, except at a very high temperature. It will not support combustion, for a lighted taper immersed in a jar of it is extinguished. It has neither colour, smell, nor taste, and in a chemical point of view it must be regarded as a very inert body, since it does not combine directly with any other substance except with great difficulty.

Nitrogen gas is erroneously called *azote*, as we are continually breathing it without perceiving any injurious effects from it. The human body is so constructed, that it will not thrive on substances intended as nourishment if they are presented to it in their purest form. Strong alcohol acts as a poison, but when diluted with four or five times its quantity of water, as in wine, it is invigorating.

Even the respiration of oxygen would soon destroy life, were it not diluted with four times its volume of nitrogen, as in atmospheric air.

Besides oxygen and nitrogen, air contains *vapour of water* and *carbonic anhydride*. The presence of the former is rendered obvious by the fall of rain, snow, dew, &c.; and that of carbonic anhydride can easily be shown by letting lime-water remain exposed to the air, or by shaking it in a flask containing air. This occasions a cloudiness in the liquid. If you ask, What is the source of this carbonic anhydride? the reply is, It is formed wherever substances are burning, wherever men and animals are breathing, and wherever decay and putrefaction are taking place. (Page 129.)

The following is a close approximation to the average composition of the air:

	By Volume.	By Weight.
Nitrogen	79	77
Oxygen	21	23
	100	100

The carbonic anhydride amounts in pure country air to about three volumes in 10,000. In towns it is more abundant.

The air is not a *compound*, but only a *mixture*. It varies slightly in composition, which compounds never do, and its properties are intermediate between those of its constituents, whereas those of a compound are entirely different.

In crowded rooms, and other confined places, the air becomes deteriorated; that is, poorer in oxygen and richer in carbonic anhydride.

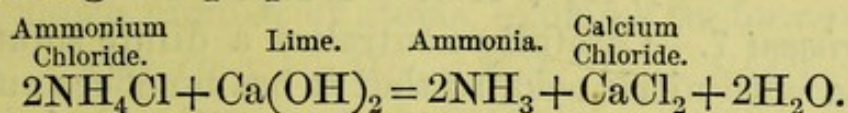
That the air also contains other foreign ingredients is not strange, since it is the constant receptacle of volatile substances and dust. The air coming from the Spice Islands, even at the distance of eight or ten miles, is impregnated with the odour of cinnamon and cloves. The dust contained in the air can be discerned in the sun-beam, &c. Minute quantities of ammonia and nitric acid are also found in the air. The presence of ozone has

already been noticed. The specific gravity of dry air, at standard pressure and temperature, is 14.47, that is, it is about $14\frac{1}{2}$ times heavier than hydrogen.

AMMONIA, NH_3 .

This important gas cannot be prepared by the direct union of its elements. It is frequently formed during the putrefaction of organic bodies, but its most convenient source is in the "ammoniacal liquor" of gas-works. From this a white substance called *sal-ammoniac*—the ammonium chloride of chemists—is manufactured.

Experiment 1.—Mix about half an ounce of powdered sal-ammoniac with about one ounce of slaked lime. A pungent odour of hartshorn is perceived. Introduce the mixture into a dry Florence flask fitted with a cork and a short tube which communicates by an india-rubber tube with a long straight tube. Apply a gentle heat, and ammonia-gas will come off in abundance. It is very soluble in water, but being lighter than air it may be collected by *upward displacement*, that is, by holding a bottle mouth downwards over the tube and allowing the light gas to displace the heavier air. The reaction by which the gas is prepared is as follows:—



Calcium chloride remains in the flask, together with the excess of lime.

Experiment 2.—Plunge a taper into a bottle of ammonia held with its mouth downwards. The taper will be extinguished and will not ignite the gas.

Experiment 3.—We must not too hastily conclude from the preceding experiment that ammonia cannot be made to burn. Hold the end of the tube from which the gas is issuing in a flame. You will perceive that the gas burns with a greenish flame *as long as the jet is heated*, but that it goes out when the lamp is removed. *Ammonia will burn when strongly heated.* During its combustion its hydrogen is oxidized to water, and its nitrogen escapes free.

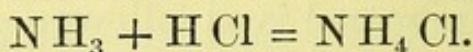
Experiment 4.—Remove the stopper downwards from a bottle of ammonia, and *afterwards* immerse the mouth of

the bottle in water. The water will rise rapidly, and if the bottle was quite full of gas, will fill it. If a piece of red litmus-paper is dipped in the water of the bottle its blue colour will return. *Ammonia is soluble in water, and its solution is a base.*

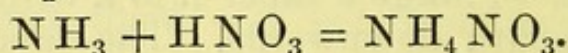
One volume of water at freezing-point will dissolve 1149, and at ordinary temperatures about 800 volumes of the gas, so that it is even more soluble than hydrochloric acid. A concentrated solution of it may be prepared like that of hydrochloric acid. It is sold under the name of solution of ammonia, or "liquor ammoniæ," and has a specific gravity of 0.88. Weaker solutions are sometimes called "hartshorn," or "spirits of hartshorn," because they were formerly obtained by the distillation of horn.

Experiment 5.—Ammonia-gas, like hydrochloric acid, can easily be prepared by gently heating its concentrated solution. By prolonged boiling all trace of ammonia may be removed from water.

Experiment 6.—The direct combination of ammonia and hydrochloric acid gases has already been noticed (page 14). The experiment may be repeated with equal volumes of the pure gases. Sal-ammoniac or ammonium chloride is formed :



Experiment 7.—Carefully neutralize a dilute solution of ammonia with nitric acid, so that it has no effect on either blue or red litmus. The solution will contain a salt called ammonium nitrate, which may be obtained in the dry state by evaporation :

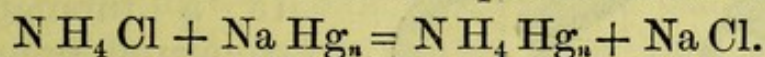


Ammonium compounds.—The compounds of ammonia are so similar in many of their properties to those of potassium and sodium, that they are generally supposed to be analogous in structure, and to contain a sort of compound metal called *ammonium*, NH_4 , which has not been isolated. When the gas dissolves in water it is assumed that the hydrate of this *quasi* metal is formed : $\text{NH}_3 + \text{H}_2\text{O} = \text{NH}_4\text{HO}$. The theory is convenient, but is inconsistent with some known facts. The following table shows the analogy which on this view exists between potassium and ammonium salts :—

K O H	Potassium Hydrate.	N H ₄ O H	Ammonium Hydrate (?)
K Cl	„ Chloride.	N H ₄ Cl	„ Chloride.
K N O ₃	„ Nitrate.	N H ₄ N O ₃	„ Nitrate,
K ₂ S O ₄	„ Sulphate.	(N H ₄) ₂ S O ₄	„ Sulphate.

It will be seen that to represent the composition of any ammonium salt we need only take the formula of the corresponding potassium or sodium compound and write NH₄ instead of K or Na.

Experiment 8.—Put about an ounce of mercury into a mortar and press into it with the pestle a few fragments of clean dry sodium. The two metals will combine with a flash of light, and a semi-fluid mass called *sodium amalgam* will be obtained. When cold, throw this into a tumbler about one-quarter full of a cold saturated solution of sal-ammoniac. The amalgam will immediately begin to swell, and will soon assume such an enormous volume as to float in the liquid as a metallic mass, which feels like butter. From the moment of its formation, however, this curious substance begins to change into ammonia, hydrogen and mercury. It is called *ammonium amalgam*, and it is believed to be formed by the temporary combination of the ammonium, which has lost its chlorine, with mercury. The following formula, in which *n* denotes an unknown number of atoms, explains its formation:



By washing with water the mercury is afterwards recovered without loss.

The specific gravity of ammonia-gas is 8.5, a little more than half that of air (14.47); at -34°C . (-29°F .) it becomes liquid, and at -75°C . (-103°F .) it freezes to a transparent, ice-like solid.

It must be borne in mind that the name ammonia is often given to the solution, which is commonly called ammonium hydrate.

COMPOUNDS OF NITROGEN AND OXYGEN.

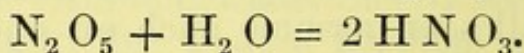
Oxides.		Corresponding Acids.
N ₂ O	Nitrogen monoxide, or Nitrous oxide	H N O Hyponitrous acid.
N ₂ O ₂ or N O	Nitrogen dioxide, or Nitric oxide	

COMPOUNDS OF NITROGEN AND OXYGEN—*continued*.

Oxides.	Corresponding Acids.
N_2O_3 Nitrogen trioxide, or Nitrous anhydride	HNO_2 Nitrous acid.
N_2O_4 or NO_2 Nitrogen tetroxide, or Nitric peroxide	
N_2O_5 Nitrogen pentoxide, or Nitric anhydride	HNO_3 Nitric acid.

Of these compounds nitric acid is undoubtedly the most important, and as, moreover, it is the source from whence the other oxides are obtained, it claims our first attention.

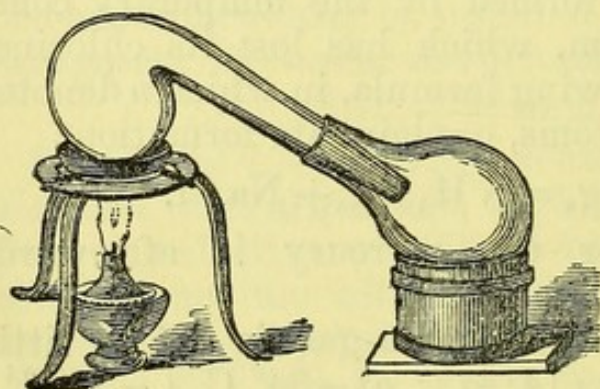
The oxide N_2O_5 , nitric anhydride, corresponding to nitric acid, is prepared with difficulty. It is a white crystalline solid, which combines eagerly with water:



Nitric Acid or Aquafortis, HNO_3 .

Experiment 1.—Introduce into a small retort half an

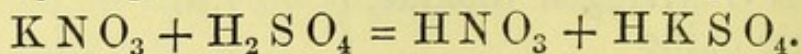
Fig. 57.



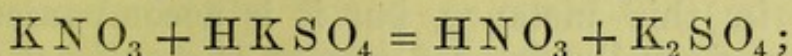
ounce of powdered salt-petre and add, by means of a thistle-headed funnel, about an equal weight of sulphuric acid. If a tubulated, i.e., stoppered, retort is used, the thistle funnel is not required. Then adapt the beak of the retort to a flask or bottle, and

apply a gentle heat. In a short time a yellowish fuming fluid passes over into the receiver, which is placed in a vessel filled with water, and must frequently be sprinkled with cold water; this fluid is heavier than water, and is called *nitric acid*.

Saltpetre is potassium nitrate, KNO_3 . When it is acted upon by sulphuric acid the metal and part of the hydrogen change places; we get nitric acid and a salt called hydrogen potassium sulphate, $HKSO_4$, in which half the hydrogen of the acid is replaced by potassium:



At a higher temperature the hydrogen potassium sulphate will decompose another molecule of potassium nitrate:



but on the small scale it is better not to push the action so far, or the retort will be apt to break. K_2SO_4 is called neutral, or di-potassium sulphate.

By using perfectly dry potassium nitrate and concentrated sulphuric acid, true nitric acid, HNO_3 , can be obtained as a yellowish liquid, which fumes strongly when exposed to the air. It is hence called *fuming* nitric acid, and is the strongest that can be prepared. It is one and a half times heavier than water, its specific gravity being 1.52. A weaker kind is commonly met with in commerce as ordinary nitric acid, or *aqua fortis*. It consists of 60 parts of true nitric acid and 40 of water, and has a specific gravity of 1.42. This acid is colourless when pure, but usually possesses a yellowish tint. When very strong or very weak nitric acid is heated, acid of this strength is obtained. It can be distilled unchanged.

Experiment 2.—A drop of nitric acid is sufficient to acidify several spoonfuls of water, and even at a greater dilution it will redden blue litmus-paper: nitric acid is accordingly distinctly characterized as an *acid*.

Experiment 3.—If lead be heated for a long time in the air it abstracts oxygen from it, and becomes converted into a reddish-yellow powder, called lead oxide, or litharge. Take up a small portion of this litharge on the point of a knife, put it into a test-tube, and add some dilute nitric acid. The greater part will be dissolved by gentle heating. Filter the solution while warm, and put it in a cold place; a salt will be deposited from it in white brilliant crystals; this is *lead nitrate*. This shows that lead oxide is a basic oxide, as it combines with acids forming salts.

Nitric acid dissolves most of the metallic oxides, and forms with them salts, all of which are soluble in water. For this reason, nitric acid is often used for cleaning metals, for instance, copper and brass instruments, which, during the process of annealing, soldering, &c., have become covered with a coating of oxide.

Experiment 4.—Pour over some shot, common nitric acid, slightly diluted with water; a solution is also effected in this instance, but it is accompanied by the evolution of a yellowish-red vapour of a suffocating smell. This vapour is nitric peroxide. Part of the nitric acid is decomposed, while another part of it combines with the lead, and forms the same salt, as in the former experiment. This likewise crystallizes from its solution, if it is evaporated until a film forms on its surface.

In this case the lead is apparently dissolved, but it is obvious that this is quite a different kind of solution from that of common salt or sugar in water. The salt and sugar are unchanged in the solution, while the lead is not contained in the liquid as a metal, but as a salt, a nitrate. The same thing occurs with all other metals which are soluble in nitric acid; as, for example, with silver, mercury, copper, iron, &c. Gold is not dissolved by it; hence gold may be separated from silver by means of nitric acid.

Nitric acid readily parts with a portion of its oxygen. It is therefore a powerful *oxidizing agent*.

Experiment 5.—Place a few fragments of tin in a glass, and pour over them nitric acid with a few drops of water. A violent action is set up, and red suffocating fumes are copiously evolved. The tin is not dissolved and converted into nitrate as the lead was, but becomes a white insoluble powder, an oxide of tin called *meta-stannic acid*.

Experiment 6.—Some of the non-metallic elements, as well as of the metals, are oxidized by nitric acid; charcoal, on being boiled in it, becomes carbonic acid; sulphur, sulphuric acid; phosphorus, phosphoric acid, &c. In all these cases yellowish-red fumes are evolved.

Experiment 7.—Organic substances also, as wool, feathers, wood, indigo, &c., are oxidized and decomposed by heating them with nitric acid. This sort of decomposition may be regarded as combustion in the moist way. If substances of animal origin are allowed to remain for a short time only in contact with this acid, they will assume a yellow colour. In this manner wood may be stained, and silk may be dyed yellow; the hands and clothes are also stained yellow by nitric acid. Cotton undergoes a most

remarkable change if soaked for a short time in the *strongest* nitric acid; it will then detonate and explode like gunpowder, only far more violently (gun-cotton). Strong nitric acid is partially decomposed, and coloured yellow, by the rays of the sun.

If you colour some water blue in a test-tube with one drop of solution of sulphate of indigo, and add to it on boiling one drop of nitric acid, the blue colour will disappear. This reaction often serves for the detection of nitric acid.

Experiment 8.—The nitrates are easily decomposed by heat, being in most cases converted into metallic oxides or metals. Having powdered some of the lead nitrate, obtained in Experiment 3 or 4, throw it upon a red-hot coal; active combustion of the coal will ensue, at the expense of the nitrate, and globules of metallic lead will remain beyond. If the lead nitrate is heated alone on a crucible cover or fragment of a broken Florence flask, yellow lead oxide will remain.

Nitric acid is *monobasic*; that is, it contains but one atom of hydrogen which can be replaced by metals. Its salts are called nitrates, and it must be remembered that they all contain the monad radical, NO_3 .

HNO_3 Nitric acid.

KNO_3 Potassium nitrate.

NH_4NO_3 Ammonium nitrate.

$\text{Cu}(\text{NO}_3)_2$ Copper nitrate.

$\text{Bi}(\text{NO}_3)_3$ Bismuth nitrate.

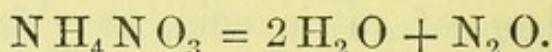
The nitrates are less easily decomposed than the chlorates, but they resemble the latter salts in the readiness with which they part with their oxygen to combustibles. In consequence of this property, potassium nitrate is a constituent of some inflammable and explosive mixtures, such as gunpowder, fuses, coloured fires, &c.

Nitrogen Monoxide, or Nitrous Oxide, N_2O .

Experiment 1.—Cautiously heat the ammonium nitrate, made by Experiment 7 (page 158), in a Florence flask to which a gas delivery tube is connected. The salt will gradually melt, and soon effervesce briskly, as a gas—

nitrous oxide—is evolved. When the air has been expelled from the flask the issuing gas may be collected at the pneumatic trough, which must be filled with warm water instead of cold, as the gas is soluble in the latter.

The change which occurs is simple. Heat decomposes ammonium nitrate NH_4NO_3 entirely into water and nitrous oxide :

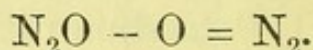


Nitrous oxide is seen to be a colourless gas. It can, however, be condensed to the liquid state by great cold.

It is respirable in moderate quantity, and its effects when inhaled are peculiar. At first it produces a pleasurable kind of intoxication—whence its name of *laughing gas*; but after a time it causes complete insensibility to pain. It is now much used as a substitute for chloroform in minor surgical operations.

Nitrous oxide is a good supporter of combustion, its power in this respect being little inferior to that of oxygen.

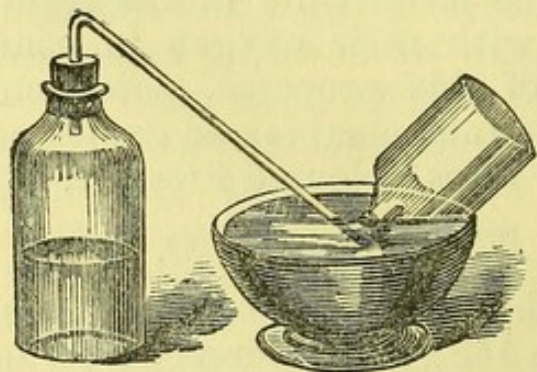
The experiments which were performed with the latter gas may be repeated with nitrous oxide, with little difference in their results. Phosphorus, for example, will burn in it with brilliancy, combining with the oxygen and leaving nitrogen behind, which is equal in volume to the original gas.



Nitrogen Dioxide, or Nitric Oxide, NO, or N₂O₂.

Experiment 1.—Pour over a few scraps of copper, placed

Fig. 58.

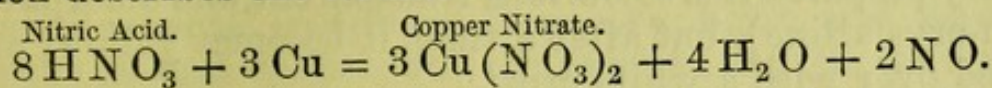


in a wide-mouthed bottle, a little water, and then add by degrees some nitric acid, until a brisk effervescence ensues. This effervescence is caused by the evolution of nitric oxide, which must be collected in a jar of white glass over the pneumatic trough. Close the mouth of the jar under water ;

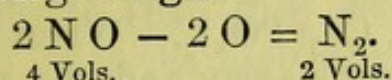
it seems to be empty, for the nitric oxide is colourless; but if the jar be opened, and air be carefully blown

in, then the jar becomes filled from above downwards with yellowish-red vapours. The nitric oxide takes thereby from the air one atom of oxygen, and is converted into nitric peroxide. NO becomes NO_2 . On account of this property, it has an important application in the preparation of sulphuric acid.

The following is the somewhat complicated equation which describes the formation of nitric oxide :

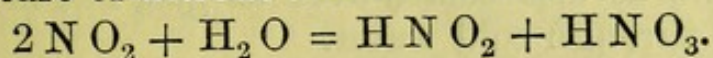


The copper nitrate remains as a blue solution in the bottle, and it may be obtained in crystals by evaporation. Nitric oxide supports combustion, although not with the readiness of nitrous oxide. If a piece of phosphorus be ignited and plunged into nitric oxide while feebly burning, it will be extinguished, but if the phosphorus burns briskly when immersed in the gas the combustion proceeds with energy. As in the previous case, the oxygen is removed by the phosphorus; the nitrogen which remains having *half* the volume of the original gas.



Nitrogen trioxide or nitrous anhydride, N_2O_3 , is a red gas, easily condensed to a blue liquid by cold. With water it forms the very unstable nitrous acid, HNO_2 .

Nitrogen tetroxide, or nitric peroxide, NO_2 , or N_2O_4 , is the red gas formed when nitrous oxide comes in contact with excess of oxygen. It is absorbed by water and converted into a mixture of nitrous and nitric acids :



PHOSPHORUS.

Symbol, $\text{P} = 31$. Formula for vapour, P_4 .

Great care is required in experimenting with phosphorus that it does not take fire at an unseasonable moment, as it continues burning with the greatest violence, and might occasion dangerous wounds. It may catch fire even when lying upon blotting-paper, particularly in summer-time, or by the heat of the finger. Hence it must be kept, and also cut, under water. On being taken from the water, it should be held by a pair of forceps, or be stuck on the

point of a knife. Prudence also would dictate to experiment with small quantities only at a time, and to have a vessel of water in readiness, in which it may be quenched in case it should catch fire.

Phosphorus, like sulphur, melts, boils, evaporates, and burns, but far more easily and rapidly. In winter it is brittle, in summer flexible as wax. When pure and freshly prepared it is colourless, transparent, and amorphous (non-crystalline), but after a time it becomes yellow, and coated over with a white crust.

Phosphorus is insoluble in water, but soluble in ether, carbon disulphide, and oils.

Phosphorus is an exceedingly violent *poison*, and is, for this reason, frequently employed for the extirpation of rats and other vermin. The rat paste, as it is called (phosphorus paste), is composed of one drachm of phosphorus, 8 ounces of hot water, and 8 ounces of flour.

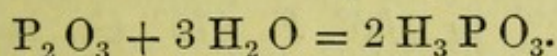
Preparation of Phosphorus.—Phosphorus was formerly obtained from urine, but is now prepared from bone-ash. Bones consist chiefly of gelatin (or rather ossein) and calcium phosphate, $\text{Ca}_3(\text{PO}_4)_2$. When the bones are calcined the gelatin burns away and the phosphate remains. The ash is treated with sulphuric acid, which converts the greater part of the calcium into the insoluble calcium sulphate, CaSO_4 , while impure phosphoric acid remains in solution. This is dried with charcoal powder and intensely heated in a clay retort. The carbon takes oxygen from the acid, forming carbonic oxide gas, CO , which escapes. The phosphorus also comes over as gas, but is condensed by passing through cold water.

The changes which occur in the above processes are somewhat more complicated than the description, which is only intended as an approximation, would imply.

Experiments with Phosphorus.

Experiment 1.—Put into a small flask first a quarter of an ounce of ether, then a piece of phosphorus, of the size of a pea. Cork the flask and let it stand some days, frequently shaking it. Decant the liquid; it contains in solution a small quantity of phosphorus, and will serve for the following experiments.

Experiment 2.—Pour some drops of this solution upon the hand, and rub them quickly together; the ether will evaporate in a few moments, but the phosphorus will remain upon the hands in a state of minutest division. The more finely it is divided, so much the more easily does it combine with the oxygen of the air. During this combination it diffuses a white smoke and a faint light (it phosphoresces), causing the hands to shine in the dark; hence its name, *phosphorus*, from $\phi\omega\varsigma$, light, and $\phi\acute{\epsilon}\rho\omega$, to carry. It undergoes slow combustion and is converted into *phosphorous anhydride*, P_2O_3 , which rapidly takes up water from the air and becomes phosphorous acid:



Experiment 3.—Moisten a lump of sugar with the solution of phosphorus in ether and throw it into hot water. The surface of the water will glow prettily in a dark room.

Experiment 4.—Instead of ether use a few drops of carbon disulphide to dissolve the phosphorus. This liquid takes up more phosphorus than ether does, and the solution is highly dangerous. It is the basis of the modern "Greek fire." Its inflammability may be observed in the following way:

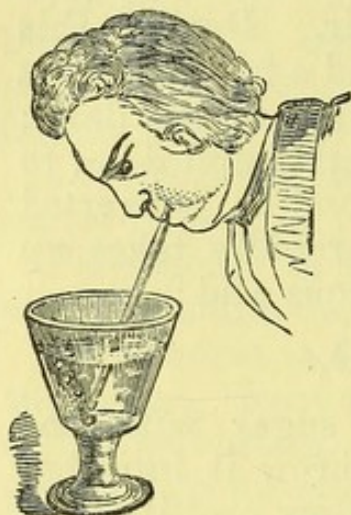
Experiment 5.—Pour some of the solution upon fine blotting-paper: the latter *ignites spontaneously* after the liquid has evaporated. The more minutely the phosphorus is divided, so much the more readily it begins to burn.

Experiment 6.—Put a piece of phosphorus of the size of a pea on blotting-paper, and sprinkle over it some soot or pulverized charcoal: it melts after a while, and spontaneously *inflames*. The finely pulverized charcoal causes this combustion, owing to its porosity. It eagerly absorbs oxygen from the air, imparts it again to the phosphorus, and, being a bad conductor of heat, the cooling of the latter is prevented.

Phosphorus is also *easily ignited by friction*, and is, for this reason, employed in the manufacture of lucifer-matches. The combustible mass is prepared from hot mucilage, at $158^{\circ} F.$ ($70^{\circ} C.$), to which small pieces of

phosphorus are added, being thoroughly incorporated with it by constant rubbing till cold. But as the mass, by becoming hard on drying, prevents the access of air to

Fig. 59.



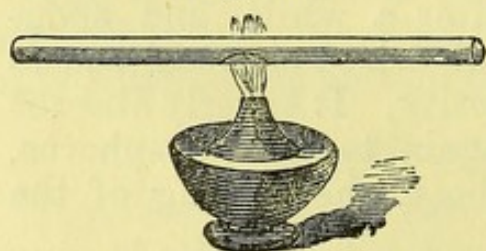
the phosphorus, there must be added some substance rich in oxygen, as black oxide of manganese, nitre, or red-lead, from which the phosphorus can abstract the oxygen necessary for its ignition. $1\frac{3}{4}$ part of phosphorus, 4 of gum arabic, 4 of water, 2 of nitre, and 2 of red-lead, form a good inflammable mass. A temperature of 149° F. (65° C.) to 158° F. (70° C.) is requisite for kindling such matches; in this case the temperature is caused by friction. The coating of the match is thus broken and kindled, and the continued burning

is now maintained by the oxygen of the air.

Experiment 7.—Put a piece of phosphorus, of the size of a pea, into a wine-glass, and pour hot water upon it until the glass is half filled; the phosphorus melts, but does not ignite, as access of air is prevented by the water. But if air be carefully blown by the mouth through a long glass tube upon the bottom of the wine-glass, *combustion* will ensue, which is visible, especially in the dark.

Experiment 8.—Heat gently a piece of phosphorus of the size of a pea, placed in the middle of a glass tube, about twelve inches long. When ignition commences, remove the lamp. While the tube is held horizontally, the

Fig. 60.



combustion is feeble and imperfect, because the heavy smoke, consisting of phosphoric and phosphorous anhydrides, passing off slowly, allows the admission of only a small quantity of air. But the combustion immediately becomes more vivid

on inclining the tube, and when the tube is held perpendicularly it is complete, as then the draught of air is most powerful. In this way phosphorus may be oxidized to either degree required; it may be slowly burnt to

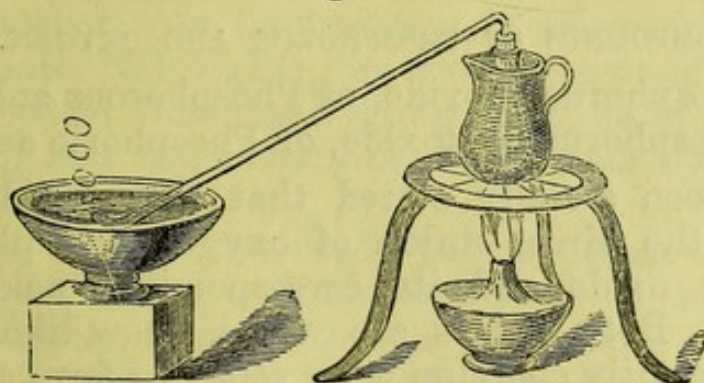
form phosphorous anhydride, or completely, to form phosphoric anhydride.

Phosphorus, like sulphur, is capable of existing in several different states (*allotropic states*). Besides the ordinary kind the most interesting is that called red, or *amorphous* phosphorus. This is prepared by heating ordinary phosphorus to a temperature of about 240°C . (464°F .) for some time, in a vessel filled with carbonic anhydride or nitrogen to prevent it from burning. Red phosphorus is curiously different in properties from the ordinary kind, being insoluble in carbon disulphide, not poisonous, and only inflammable at a high temperature. It is now used extensively in the manufacture of "safety matches." The tips of these matches do not contain phosphorus, but when rubbed on the outside of the box they take up a little of the red phosphorus with which the side is covered, and combustion ensues.

PHOSPHINE, PH_3 .

Experiment 1.—Put into an ounce flask a quarter of an ounce of slaked lime, and a piece of phosphorus the size of a pea, fill it up to the neck with water, and place it in a small vessel containing a strong solution of salt, prepared

Fig. 61.



by adding half an ounce of common salt to an ounce and a half of water. Fit to the flask a bent glass tube, one end of which is made to dip into a basin of water; heat the salt water to boiling, and a gas will be evolved, which, as it issues from the tube and comes in contact with the air, *takes fire spontaneously*. This gas consists chiefly of phosphine, or phosphuretted hydrogen. If you collect it in a small jar filled with water, it immediately

takes fire upon the admission of air. Both the phosphorus and the hydrogen combine with the oxygen of the air, and there results phosphoric anhydride, P_2O_5 , and water, H_2O . The first appears as a white smoke, which, when it issues in separate bubbles from the water, rises in rings.

Phosphine, when unburnt, emits the smell of garlic.

In this experiment, the flask is placed in salt water, in order to guard against the ignition of the phosphorus, in case the flask should accidentally break. Salt water, at the strength specified, will not boil under $228^{\circ}F.$ ($109^{\circ}C.$); consequently, the boiling in the flask is more active than if it had been placed in pure water, the temperature of which, under ordinary pressure, can only be raised to $212^{\circ}F.$ ($100^{\circ}C.$). The apparatus for heating substances by means of hot water or saline solutions is called a *water* or *saline bath*. By such contrivances extracts are evaporated, and substances dried, which, at a stronger heat, would easily burn, or otherwise be decomposed.

Phosphine, when pure, is inflammable, but *not spontaneously inflammable*. When prepared as above it contains a little of the vapour of another compound of phosphorus and hydrogen—a liquid—which has the formula P_2H_4 . It is the latter compound which is spontaneously inflammable.

COMPOUNDS OF PHOSPHORUS AND OXYGEN.

P_2O_3 , Phosphorus trioxide, or Phosphorous anhydride.

P_2O_5 , Phosphorus pentoxide, or Phosphoric anhydride.

It has been already stated that these oxides can be formed by the direct union of oxygen and phosphorus. The lower oxide and its corresponding acid are unimportant. Both oxides are white, snow-like powders, and they greedily combine with water to form acids.

Phosphoric Anhydride.

Experiment 1.—Ignite a small piece of phosphorus on a plate and cover it with a *dry* bottle or bell-jar. Phosphoric anhydride will be abundantly formed, as a white cloud, which settles on the sides of the bottle and on the plate. Shake the powder into a saucer, and pour a little water on it. A sharp hissing noise will be heard, testifying to the

energy with which combination ensues. A solution of phosphoric acid is produced.

Phosphoric Acid, H_3PO_4 .

Experiment 2.—Place a piece of phosphorus and a little water in a flask, and carefully heat them until the water boils; then add a little nitric acid and continue the heat. The phosphorus slowly dissolves, and is oxidized into phosphoric acid. Phosphoric acid prepared in this way is of course contaminated with the excess of nitric acid left undecomposed. This can be removed by evaporating to dryness and boiling the residue with water.

Phosphoric acid is tribasic, containing three replaceable atoms of hydrogen; and inasmuch as one, two, or all three of these atoms may be replaced by metals, its salts are very numerous. With sodium, for instance, the three following salts may be obtained:

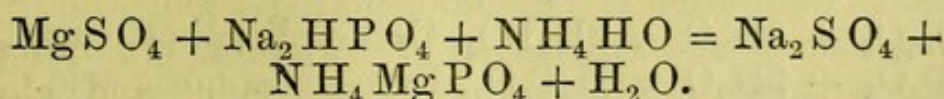
NaH_2PO_4 , Sodium di-hydrogen phosphate.

Na_2HPO_4 , Di-sodium hydrogen phosphate.

Na_3PO_4 , Tri-sodium phosphate.

A very great variety of compound salts may be formed by substituting part of the hydrogen with one metal, and part with another. As, for instance, in ammonium magnesium phosphate, NH_4MgPO_4 , in which one atom of hydrogen is replaced by the monad group of atoms, or compound radical, NH_4 , and the remaining two atoms by the diad metal magnesium, Mg.

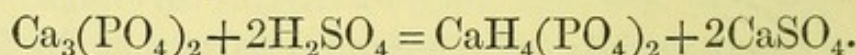
Experiment 3.—Add ammonia to a solution of sodium phosphate, and then a little solution of magnesium sulphate, $MgSO_4$ (Epsom salts). A white precipitate will be produced, consisting of the above-mentioned ammonium magnesium phosphate:



Phosphoric acid or any other soluble phosphate would produce the same effect. Ammonium hydrate and magnesium sulphate therefore serve as a test for phosphoric acid and its salts.

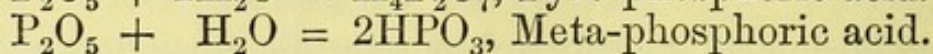
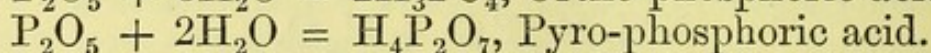
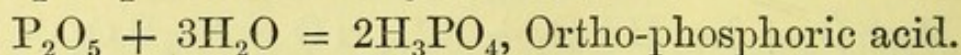
The body of an adult contains
 from 9 to 12 pounds of bones, containing
 „ 6 „ 8 pounds of bone ashes, containing
 „ 5 „ 7 pounds of calcium phosphate, containing
 „ 1 „ $1\frac{3}{4}$ pounds of phosphorus.

Phosphates are also contained in the blood, flesh and other portions of the body. Whence does it obtain this phosphorus? The answer is, from the meat and vegetables which it consumes. The phosphates occur in bread, in all kinds of grain, in leguminous and many other plants, particularly in their seeds. But how do the plants obtain these salts? By means of the soil. If arable land contained no such salts, no seeds could be produced. If we increase their quantity by mixing ground bones with the soil, we place the latter in a situation to produce a larger quantity of grain; consequently, bones furnish us with a powerful manure. The tri-calcium phosphate present in bones is however insoluble in water, and is absorbed but slowly by soils. It is therefore generally treated with sulphuric acid, by which a mixture of calcium sulphate and acid calcium phosphate, both soluble in water, is obtained.



The mixture is largely used as manure, and is known as superphosphate of lime.

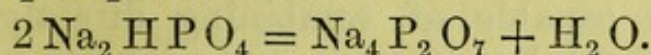
Besides ordinary phosphoric acid, two other varieties of phosphoric acid are known, differing in constitution from ordinary phosphoric acid by the relative quantities of water combined with the anhydride. These compounds are distinguished by the prefixes “pyro” and “meta” from ordinary or *ortho*-phosphoric acid. To form *ortho*-phosphoric acid three molecules of water are necessary, while for *pyro*-phosphoric acid two are required, and for *meta*-phosphoric acid only one:



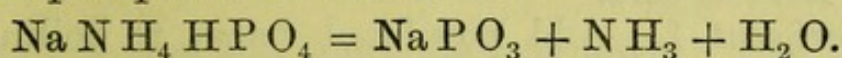
These three acids form salts which correspond to them in constitution. The common or *ortho*-phosphates contain the triad radical PO_4 , the *pyro*-phosphates, the tetrad

radical P_2O_7 , and the meta-phosphates, which resemble the nitrates in constitution, the monad radical, PO_3 .

Experiment 4.—Heat common sodium phosphate to redness in a crucible for some time. Water is driven off, and sodium pyro-phosphate remains :

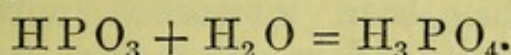


Sodium meta-phosphate may be prepared by heating microcosmic salt, $NaNH_4HPO_4$ (sodium ammonium hydrogen phosphate). Ammonia and water are expelled, and the meta-phosphate remains :



These pyro- and meta-phosphates may be dissolved in cold water without change, but if the solutions are boiled water is taken up and common phosphates obtained.

When common phosphoric acid is evaporated to dryness, meta-phosphoric, or “glacial” phosphoric acid is obtained. This, when boiled with water, is reconverted to ortho-phosphoric acid :



Experiment 5.—To solutions of sodium ortho-, pyro-, and meta-phosphate add a little solution of silver nitrate. With the ortho-phosphate, a yellow precipitate, and with the others white precipitates are produced. This test distinguishes the ortho-phosphate from the other two. To distinguish between these, add to each of the salts, first some acetic acid, and then a little white of egg beaten up with water. A white precipitate is produced with the meta-phosphate ; none at all with the pyro-phosphate, or ortho-phosphate.

BORON.

$$B = 11.$$

Boron is a triad element, but it resembles carbon in its physical characters. Like carbon it can be obtained as an amorphous powder, and also in the form of octahedral crystals which are nearly as hard as diamonds. Boron is chiefly found combined as boric acid in the steam which issues from fissures in the earth in some parts of Tuscany, and bubbles up through small lakes or basins of water. The

acid is obtained by evaporating the water of these lakes by an ingenious arrangement, whereby the heat required for the purpose is supplied by the hot springs themselves.

A sodium salt of boric acid called *tincal* is imported from Thibet; when purified by recrystallization, it is sold as *borax*.

Boron is capable of combining directly with many elements, as sulphur, oxygen, and nitrogen.

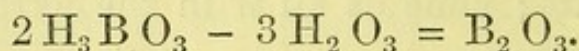
The oxygen compound of boron is the most important, the others are of little interest.

Boric anhydride, B_2O_3 , and *Boric acid*, H_3BO_3 , sometimes called *Boracic acid*.

Experiment 1.—Add powdered borax to boiling water until it ceases to be dissolved, pour off the solution from any undissolved borax, and mix it with about half its bulk of strong hydrochloric acid. Flaky crystals are immediately formed, and deposit abundantly as the mixture cools. Collect the crystals in a filter, wash them with a little cold water, preserve, and label them boric acid, H_3BO_3 .

Experiment 2.—Pour alcohol on some boric acid, or on borax which has been moistened with oil of vitriol. In flame the spirit, and in either case it will burn with a greenish flame, from the presence of boric acid. Curiously enough, *dry* boric acid may be heated to redness without volatilizing, but if it be boiled with water or alcohol it readily passes off in the vapour of those substances.

Experiment 3.—Heat boric acid in a crucible with a Bunsen's burner. Below a red heat it fuses, and on cooling appears as a colourless glassy mass consisting of boric anhydride, B_2O_3 :



Experiment 4.—If boric acid is heated at a temperature of about $120^\circ C.$ ($248^\circ F.$) part only of the water is driven off, and a substance remains called meta-boric acid, HBO_2 .

Borax is an irregular substance, having the formula $Na_2B_4O_7$. It may be regarded as a compound of sodium metaborate with boric anhydride, $2NaBO_2, B_2O_3$. Borax is often employed in metallurgical operations as a flux. It is also used in blowpipe experiments, because some metallic oxides give coloured beads when fused with it on a loop of platinum wire.

CHAPTER IV.

NON-METALLIC TETRAIDS.

(Carbon, Silicon.)

CARBON.

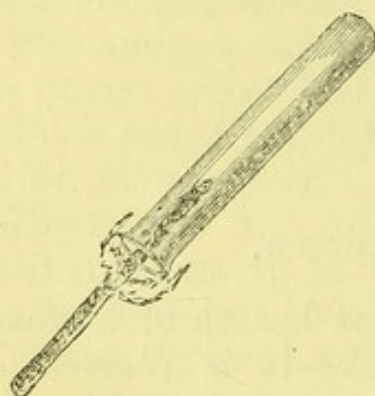
C = 12. Formula unknown.

If a piece of wood be placed on the hot hearth of a stove, it becomes brown, and finally black—it is *charred or carbonized*. If water be poured upon a burning chip, the latter is extinguished—it is seen to be also *carbonized*. A piece of linen, when inflamed and immediately smothered, becomes tinder. Tinder is *carbonized* linen. In the first case, the heat was not sufficiently strong entirely to burn the wood; in the second, the complete combustion was prevented by cooling; and in the third, by the exclusion of air. *Many animal and vegetable substances if only partially burnt yield charcoal*. As charcoal, on exclusion of the air, cannot be melted, even in the strongest heat, so its external appearance is very different, according to the nature and structure of the substance from which it was prepared; indeed, this difference often extends to the properties of charcoal, as in wood-charcoal, soot, coke, bone-black, &c. The black matter of the different kinds of charcoal consists of the element CARBON. Organic bodies always contain carbon. When they are charred, a great part of the other constituents is driven off, together with a portion of the carbon, as is shown by the fact that the carbonized body weighs much less than the original substance.

Carbon also exists in the mineral kingdom. It forms the principal element of common coal, brown coal, &c., which have all been formed from the vegetation of an earlier period, and also of the mineral oil and tar which occur in such large quantities in America and other

parts of the world. It is found almost pure as diamond and graphite (black lead), and, combined with oxygen, is contained in limestone, marble, chalk, and various other minerals.

Charcoal.—*Experiment 1.*—Gradually introduce a burning splinter of wood into a test-tube. *Fig. 62.*



ing splinter of wood into a test-tube. The part outside of the tube only will burn with a flame, while that within merely carbonizes, because the air is excluded. On the same principle, wood-charcoal is prepared on a large scale. Piles of wood (charcoal kilns) are erected, which are covered with turf and moist earth, and the wood is then kindled. This would be

extinguished, however, for want of air, if holes were not made, by wooden pokers, at different parts of the kiln, through which fresh air may be admitted, and the burnt air may escape. Only so much air should be admitted as is necessary for carbonizing or half-burning the wood. One pound of wood commonly yields about one quarter of a pound of charcoal. The quantity, however, varies much with the temperature and the nature of the wood.

Experiment 2.—Weigh a piece of freshly-burnt charcoal, and let it remain for a day in a moist place; it will now weigh more than before, owing to its having *imbibed air and moisture*. If the charcoal be now put into hot water, the air will escape from it in numerous bubbles, being expelled by the heavier water, which replaces the air in the small interstices or pores of the charcoal.

Experiment 3.—Cover a dead rat to a depth of about an inch with coarsely-powdered wood-charcoal. The animal will soon putrefy, but the vessel will not give out any disagreeable odour even if kept for years. All the noxious gases which are given out are absorbed and destroyed by the charcoal.

Experiment 4.—Foul stagnant water is deprived of its bad taste, and is rendered clear and colourless, by being filtered through charcoal. In some large cities, where there is a scarcity of potable water, it is not unusual to

filter through charcoal. Charcoal renders ordinary brandy pleasanter in taste and smell, by absorbing into its pores an acrid volatile oil, *fusel oil*, with which crude brandy is contaminated. It also deprives beer of its bitterness, by absorbing certain component parts of the hops.

Fig. 63.

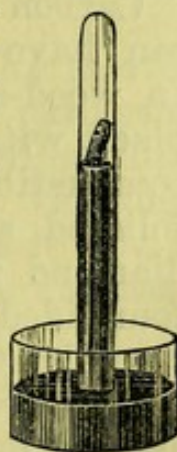


The cause of these remarkable properties of charcoal is to be found in the number of pores with which it is permeated. Into these pores, gases are absorbed, sometimes in enormous quantity. Cocoa-nut charcoal introduced into ammonia-gas standing over mercury, will absorb 171 times its volume of the gas. Now under ordinary circumstances charcoal, being exposed to the air, contains a great deal of absorbed oxygen, and it is this oxygen which destroys the noxious matters with which the charcoal comes in contact.

In addition to wood-charcoal, the following other kinds of charcoal have many different applications.

Soot, or *lamp-black*, is charcoal in a state of minute division, which is deposited from carbonaceous gases, commonly from illuminating gas; for instance, from the flame of common coal, wood, oil, rosin, &c., when during the combustion there is an insufficient supply of air. One variety of superior quality is called lamp-black. The soot must be freed from empyreumatic substances, by igniting it thoroughly in a well-closed vessel. Soot, as is well known, is our most important black colouring substance (Indian ink, printing-ink).

Fig. 64.



Coke, or carbonized coal, has a grey colour, is more or less porous, is very hard and has a metallic lustre; it burns without forming soot, and gives out an intense heat; hence it is an excellent fuel, and especially adapted for the smelting of iron, and for the heating of locomotive boilers. Coke is obtained as a secondary product in the preparation of illuminating gas from coal.

Bone-black, or *animal charcoal*, is obtained by heating bones in close vessels. The carbon contained in it amounts to little more than one-tenth part of the whole, the rest

being bone-ash ; but, notwithstanding this, its decolorizing power is so strong, that it is preferred to all other kinds of charcoal as a means of abstracting colour from the syrup of brown sugar, or from other dark liquids.

Two sorts of carbon found in the mineral kingdom, viz., graphite and the diamond, possess very remarkable, yet different, properties.

Graphite, or plumbago, a grey substance, having a metallic lustre, imparts its colour so readily to other bodies, that it is used for making pencils, and for giving a black polish to iron articles, such as stoves, &c. ; it is so soft and lubricating, that it is added to grease for the purpose of preventing friction in wheels and machinery ; it is also so nearly incombustible, that it is sometimes used in the manufacture of crucibles.

Diamond is the hardest of all bodies. In external appearance it has not, indeed, the slightest resemblance to charcoal, yet it can be entirely burnt up in oxygen, and carbonic anhydride is the only product obtained from it ; and exactly so much is obtained as would have resulted from the combustion of an equally heavy piece of purified wood-charcoal or coke. At a high temperature and under great pressure minute diamonds can be obtained artificially.

Carbon shows very clearly how one and the same body can have quite different forms and different properties. In wood-charcoal, soot, coke, and animal charcoal, it is black, without any determinate shape (*amorphous*), and very combustible ; in graphite it is black, with a crystalline foliated structure, and is nearly incombustible ; in the diamond it is colourless, and is crystallized as a double four-sided pyramid (octahedron), and is likewise almost incombustible. Hence carbon is said to be *dimorphous*, having two different crystalline forms. If a body can assume more than two crystalline forms it is said to be *polymorphous* (having many forms).

Charcoal undergoes no change on exposure to the air, or when imbedded in the ground. It is not changed at common temperatures, that is, it does not enter into combination with the oxygen of the air or of water. But this, as is well known, takes place very readily, when

heated to redness. It then burns and disappears, with the exception of a small quantity of ash. The heat thus developed is the result of its chemical combination with the oxygen of the air. The gas generated is called *carbonic anhydride*, which forms, with lime-water, a white precipitate (calcium carbonate), as has been stated previously.

COMPOUNDS OF CARBON AND OXYGEN.

Carbonic anhydride, or carbon dioxide, CO_2 . Carbonic acid H_2CO_3 (?).

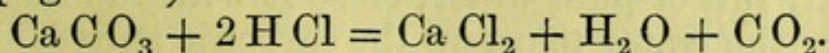
Carbonic oxide, or Carbon monoxide, CO .

Carbonic anhydride, CO_2 , sometimes called carbonic acid gas.—It has already been shown (page 128) that all ordinary kinds of fuel contain carbon, and therefore yield, during their combustion, carbonic anhydride, and that this gas may be detected by lime-water, which is thereby rendered turbid, owing to the formation of calcium carbonate.

Calcium carbonate occurs in nature in immense quantities, as chalk, marble, limestone, calc spar, &c.

By treating calcium carbonate with almost any acid it is decomposed, and carbonic anhydride is liberated, which is thus procured more conveniently than by burning carbon.

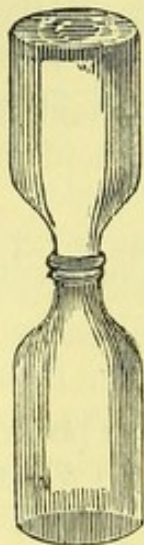
Experiment 1.—Place some fragments of white marble with a little water in the apparatus used for preparing hydrogen, and then add small quantities of hydrochloric acid until a brisk effervescence is produced. Carbonic anhydride is rapidly evolved, and as it is a heavy gas, it may be collected in dry bottles by downward displacement (page 116). The reaction is as follows:



Solution of calcium chloride remains in the bottle when all the marble has disappeared, and it may be obtained solid by evaporation.

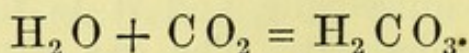
Experiment 2.—A burning taper is extinguished when held in carbonic anhydride, and it is fatal to men and animals if they inhale it. The gas can *neither support combustion nor life*.

Experiment 3.—Invert a jar filled with carbonic anhydride over one containing only atmospheric air; if after some moments you introduce into each of these jars a burning taper, that in the upper vessel will continue to burn, while that in the lower one will be extinguished. Carbonic acid is *heavier* than common air; it has sunk into the lower jar, while the atmospheric air has ascended into the upper one. If a bottle, filled with carbonic anhydride, be held with its mouth obliquely over the flame of a lamp, so that the gas can flow out, the light will be extinguished.



Experiment 4.—Repeat the experiment with the two jars, filling, instead of the upper one, the *lower* one with carbonic anhydride. If, after some hours, you add lime-water to both of the jars, and shake them, you will obtain in both of them a precipitate of calcium carbonate—a proof that the carbonic anhydride has partly ascended into the upper jar. Both gases have become intimately mixed, or the carbonic anhydride, though heavier, has ascended, and the common air, though lighter, has diffused itself downwards. This voluntary mixing of the different kinds of gases together is called *diffusion of gases*. This diffusion of gaseous bodies, since it maintains a constant equality and balance of the constituents of the atmosphere, is of great importance in the economy of nature, and accounts for the fact that the constitution of the air is everywhere nearly uniform, although in one place free oxygen is withdrawn from it, and in another place carbonic acid is added to it.

Experiment 5.—Fill a bottle containing carbonic anhydride half full of pure water, close it with the finger, and shake it; the water takes up the carbonic anhydride, and, as a vacuum is formed, the finger is pressed into the mouth of the bottle by the external air. Carbonic anhydride is *soluble in water*; and the solution is supposed to be one of carbonic acid:



One volume of water will absorb about one volume

of carbonic anhydride at the ordinary temperature of the air. If the gas be condensed by pressure, water will still absorb the same volume, but of course a much greater weight. The water thereby acquires an acid taste, and the property of effervescing.

Experiment 6.—Throw a piece of chalk into vinegar; vinegar is one of the weakest of common acids, yet it is able to expel carbonic anhydride.

Carbonic anhydride can be converted into a *liquid* by strong compression at a low temperature. This liquid boils at a temperature of 108°F. (-78°C.); when allowed to escape from a jet, intense cold is produced by the rapid evaporation (page 41), and a portion is frozen into a white snow-like solid. The solid, mixed with ether, will freeze mercury instantly.

Experiment 7.—To prove that carbon is contained in the colourless carbonic anhydride, take a test-tube, break the bottom of it, and adapt to it, by means of a perforated cork, a glass tube, and connect it with a bottle in which carbonic anhydride is evolved. Introduce into the test-tube a piece of potassium of the size of a pea, previously dried between blotting-paper, and heat the place where it lies with a lamp. The potassium will remove the oxygen from the CO_2 and black carbon will remain in the tube.

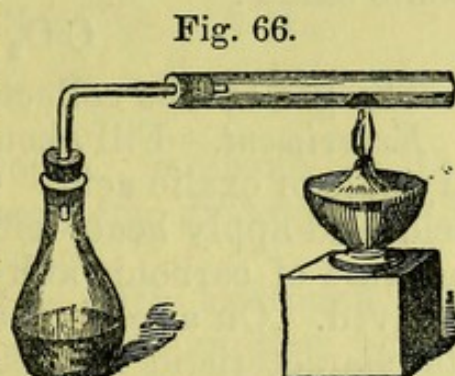


Fig. 66.

Carbonates.—Although we must regard carbonic acid, H_2CO_3 , as hypothetical, its salts are well-known and stable compounds. The following are examples of important carbonates:

NaHCO_3 , Sodium hydrogen carbonate.

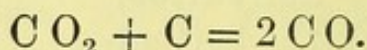
Na_2CO_3 , Sodium carbonate.

CaCO_3 , Calcium carbonate.

Carbonic oxide, CO.—When charcoal during combustion has a sufficient supply of oxygen, carbonic anhydride, CO_2 , is formed; but if there is a deficiency of oxygen, the lower oxide, CO, is produced.

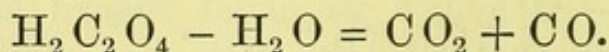
This gas, which is a deadly poison, is formed when charcoal burns slowly, for example, in a chafing-dish, and it is also formed when the damper of a stove is closed, before the coal is burnt out, since in this case the draught of air, and consequently the supply of sufficient oxygen, is prevented. Carbonic oxide burns when kindled, with a blue flame; it takes up the deficiency of oxygen not supplied to it by the air while forming, and is converted into carbonic anhydride; that is, it takes up as much oxygen again, and CO becomes CO₂. The blue flame which is often perceived in large masses of glowing coals is burning carbonic oxide gas.

Carbonic oxide can be prepared by passing carbonic anhydride through a red-hot tube containing fragments of charcoal: one molecule of carbonic anhydride combines with one atom of carbon, and forms *two* molecules of carbonic oxide:



The gas can be collected at the pneumatic trough.

Experiment.—Fill about half an inch of a test-tube with crystals of oxalic acid. Cover them with strong sulphuric acid and apply heat—effervescence will take place, and a mixture of carbonic anhydride and carbonic oxide will be evolved. On applying a light the latter will burn with its characteristic blue flame, in spite of the presence of carbonic anhydride. If the gases are passed through a strong solution of caustic soda the CO₂ will be absorbed and the CO obtained in a pure state. The sulphuric acid acts by removing water:



A great number of compounds of carbon and hydrogen are known, but most of them belong to the department of organic chemistry. A classification of them will be found in Part III., Chap. I. It will, however, be convenient to study two or three of them in this place.

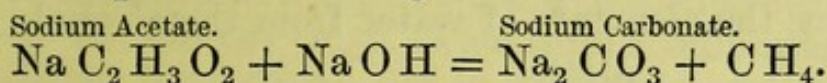
Methane or Marsh Gas, CH₄.—If the mud at the bottom of a pond is stirred up with a stick, bubbles of gas rise which consist chiefly of marsh gas, mixed with carbonic anhydride.

Experiment 1.—Fill a wide-mouthed bottle with water,

invert it in a pond and collect the bubbles which rise on stirring. The bottle can then be removed with a saucer. If a piece of moist potash is then passed up into the bottle, the carbonic anhydride will soon be absorbed. The remaining gas will burn when a flame is applied to its mouth.

This experiment proves, 1, that methane is formed during vegetable decay, and, 2, that it is an inflammable gas.

Experiment 2.—Dry some sodium acetate thoroughly at a gentle heat, and weigh out half an ounce of it. Then add half an ounce of dry caustic soda and three-quarters of an ounce of quicklime. Grind them together quickly in a mortar, and introduce the mixture into a dry Florence flask fitted with a somewhat wide delivery tube. On the application of heat methane is given off, and can be collected at the pneumatic trough:

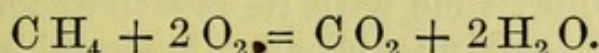


The lime is only used to prevent the action of the alkali on the glass.

Experiment 3.—Next to hydrogen, methane is the lightest of all known gases. Its specific gravity is only 8 (eight times heavier than hydrogen).

Balloons may be filled with it as with hydrogen, and the gas be transferred from one vessel to another by the method of *inverted pouring* used in the case of hydrogen.

Experiment 4.—We have already seen that methane is inflammable. In burning it is converted into carbonic anhydride and water:



Fill a soda-water bottle one-third with methane, and two-thirds with oxygen. On the application of a light a loud explosion will be produced. We see therefore that, like hydrogen, methane will burn either quietly or with explosion, according to circumstances.

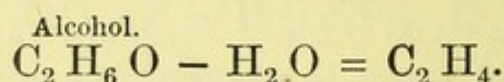
The dangerous *fire-damp*, as the miners call it, which is given out from the beds of coal in coal-mines, is almost pure methane. The disastrous explosions we so often hear of are caused by the accidental contact of a light with a mixture of fire-damp and air.

Methane is not poisonous. Miners can work in an atmosphere largely charged with it, but its combustion produces the deadly poison carbonic anhydride. This hangs about the passages of a mine after an explosion, and often suffocates the poor creatures who have escaped from the explosion. The miners call it *choke-damp*, or *after-damp*.

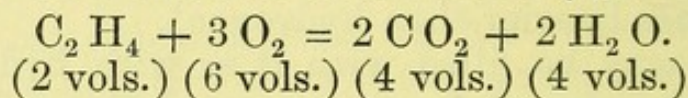
Methane is sometimes called "light carburetted hydrogen."

Ethylene, or Olefiant gas, C_2H_4 . — Sometimes called "heavy carburetted hydrogen."

Experiment 1. — Mix one volume (half an ounce *by measure*) of strong alcohol, or methylated spirit, with four volumes of strong sulphuric acid. Heat the mixture in a flask and collect over warm water the gas that is given off. It is ethylene. The sulphuric acid acts by removing the elements of water from the alcohol :

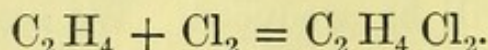


Experiment 2. — Ethylene may be burned at a jet, or at the mouth of a bottle, like methane. To explode it, it must be mixed with three volumes of oxygen :

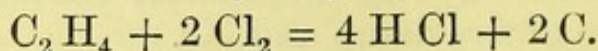


This will be apparent from the above equation, since the formula for every gas denotes two volumes of it.

Experiment 3. — Mix equal volumes of ethylene and chlorine in a bottle over water. Allow the bottle to remain for a short time with its mouth under water. The two gases will combine, and an oily liquid, called *ethylene chloride*, or *Dutch liquid*, will be produced, and will float on the water :



Experiment 4. — Mix one volume of ethylene with two volumes of chlorine in a tall bottle or jar, and apply a light. A very curious combustion, attended with the formation of thick black smoke, will take place. The chlorine combines with hydrogen, and *carbon* is separated :



Acetylene, C_2H_2 .—This interesting gas is formed by the direct union of its elements when the electric light passes between two charcoal points in an atmosphere of hydrogen.

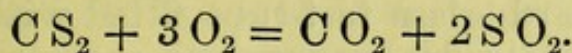
All the processes for preparing it are difficult. Methane, Ethylene and Acetylene are all constituents of coal-gas. The two last have remarkable illuminating power.

CARBON DISULPHIDE, CS_2 .

Just as charcoal at a red heat will combine with oxygen, so it will with sulphur. The compound so produced is called carbon disulphide, or "bisulphide of carbon." It is a heavy liquid (specific gravity 1.27), which boils at $46^\circ C.$ ($109^\circ F.$), has a disgusting odour, and is extremely inflammable.

Experiment 1.—Boil a little oil in a test-tube, and bring the end of the tube in contact with a few drops of carbon disulphide in an egg-cup. The liquid will ignite and burn with a blue flame. It is therefore evident that carbon disulphide ignites at a very low temperature.

Experiment 2.—Shake a few drops of the disulphide in a wide-mouthed bottle, and apply a light. The vapour will burn with a blue flame, or if the quantity of liquid was very small, a slight explosion will be produced. A portion of the sulphur escapes oxidation and remains on the sides of the bottle. When, however, the liquid or vapour burns in an open vessel or at a jet, both the sulphur and carbon are oxidized, carbonic and sulphurous anhydride being formed:



Experiment 3.—Repeat the last experiment with a half-pint wide-mouthed bottle of oxygen. A very loud explosion will be produced.

Experiment 4.—Very fine iron wires may be burned in the flame of carbon disulphide, the iron combining with the sulphur to form *ferrous sulphide*, FeS .

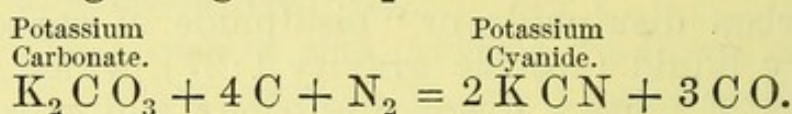
In consequence of its remarkable volatility and inflammability, carbon disulphide is a dangerous liquid to work with. All the above experiments must be made with small quantities, and with the greatest caution.

Carbon disulphide is a very useful solvent. Iodine,

sulphur, phosphorus, and fixed oils are among the substances which can be dissolved in it.

CYANOGEN, C_2N_2 or $(CN)_2$.

Carbon and nitrogen do not unite directly with one another, but by various indirect processes a series of important salts called *cyanides* can be prepared. When, for instance, a mixture of potassium carbonate and charcoal is heated in a tube, and nitrogen gas is passed through it, the following change takes place:

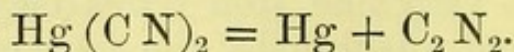


Potassium cyanide is in practice prepared from the salt called potassium ferrocyanide, or yellow prussiate of potash. The last-named salt, which is a very important article of commerce, is prepared by heating strongly a mixture of potassium carbonate, iron filings, and refuse animal matters, such as cuttings of horn and leather, woollen rags, dried blood, and the like. From potassium cyanide other cyanides can readily be prepared.

The cyanides all contain the monad radical CN, which is often denoted by the symbol Cy, just as if it were an elementary atom. The cyanides are, in fact, closely analogous to the chlorides.

Cyanogen itself, and many of the cyanides, are deadly poisons.

Experiment 1.—Heat a small quantity of mercuric cyanide in a hard glass test-tube fitted with a short jet. The white crystals decompose into mercury, which condenses on the upper part of the tube, and cyanogen gas, which escapes, and can be burned at the jet. It burns with a beautiful peach-coloured flame:



A brown substance, called *paracyanogen*, is formed at the same time.

It will be seen that cyanogen gas is the radical CN *doubled*. Its formula is therefore often written as $(CN)_2$, or Cy_2 . The resemblance of cyanogen and the cyanides to chlorine and the chlorides is then perfect:

Cyanogen	$(\text{CN})_2$	or Cy_2	like Cl_2 .
Hydrocyanic Acid.	HCF	„ HCy „	HCl .
Potassium Cyanide	KCN	„ KCy „	KCl .
Mercuric Cyanide	$\text{Hg}(\text{CN})_2$	„ HgCy_2 „	HgCl_2 .

Cyanogen is soluble in water. To obtain it pure it must be collected over mercury, for it would not be safe to collect it by displacement.

Hydrocyanic Acid, HCN , or HCy .—Often called *prussic acid*. This deadly poison is obtained by the action of sulphuric acid on cyanides. It is an exceedingly volatile liquid, soluble in water in all proportions.

Experiment 2.—Put a few grains of potassium cyanide into a small beaker, and moisten the salt with dilute sulphuric acid. Hydrocyanic acid will be set free with effervescence, and its peculiar odour—that of bitter almonds—will be recognised. It must, however, be remembered that the acid is most deadly as vapour.

Experiment 3.—Repeat the last experiment, covering the beaker with a piece of glass moistened with a drop of silver nitrate. The hydrocyanic acid vapour will decompose the silver salt, and the white insoluble *silver cyanide*, AgCy , will be formed. This substance is very like silver chloride, except that it does not turn purple on exposure to light, and is soluble in hot nitric acid. Like the chloride, it dissolves in ammonia.

Experiment 4.—Moisten the glass plate with a drop of solution of potassium hydrate (potash) and expose it for some minutes to the vapour. No change will be perceived, but the potash will be partly converted to potassium cyanide. Add a drop of ferrous sulphate (green vitriol) to the liquid on the plate, and then, after standing for ten minutes, a few drops of hydrochloric acid. A beautiful blue precipitate (Prussian blue) will be formed.

This is a most delicate test for cyanogen.

Experiment 5.—The salt called potassium ferrocyanide has the formula $\text{K}_4\text{Cy}_6\text{Fe}$. It is regarded as containing a tetrad radical (Cy_6Fe), the iron being a part of the radical. Many other ferrocyanides can be formed from this one by double decomposition. Add solution of potassium ferrocyanide to solutions of the following salts. In each case a ferrocyanide of the metal will be precipitated :

Copper Sulphate.—Purple brown.

Lead Acetate.—White.

Ferrous Sulphate.—White, or pale blue, soon changing to dark blue.

Ferric Sulphate or Chloride.—Dark blue (Prussian blue). If the ferrous salt be absolutely free from ferric salt, a white precipitate is obtained, but this is rarely the case.

DESTRUCTIVE DISTILLATION.—FLAME.

When wood, coal, and many other substances are heated in close vessels, an entire rearrangement of their constituent elements takes place. A portion of the carbon remains behind as charcoal, or coke, the rest distils over in combination with the other constituents, in the form of a number of volatile compounds, some of which readily condense into liquids or solids, while others are permanent gases. The nature of these volatile products cannot be fully discussed in this place. Many of them are combustible—these are for the most part compounds of carbon and hydrogen. Most can be applied to some useful purpose.

Processes of this kind, in which substances are *destroyed* in the retort, and volatile compounds formed and distilled from them, are called processes of *destructive distillation*. The preparation of gas from coal is the most important example.

Distillation of Coal.—*Experiment 1.*—Take a piece of hard tube about half an inch in external diameter, and a foot in length. Convert it into a long test-tube by drawing off a short piece from one end, and then bend it twice at right angles thus: *N*. Fit the open end with a small cork and short jet. This is a very useful piece of apparatus for experiments of this kind, acting at once as retort and receiver. Introduce some fragments of dry coal into the closed end and apply a strong heat (a Bunsen's burner does very well). *Water* and *tar* will distil over and condense in the bend, while *coal gas* will pass over and may be ignited at the jet. When no more gas is evolved, *coke* will remain in the retort.

Gas Manufacture.—On the large scale, the coal is

distilled in iron or clay retorts. The vapours pass first into the *hydraulic main*, a broad iron pipe placed horizontally in front of the retorts. From this it passes through a series of pipes called the *condensers*. In the hydraulic main and condensers the gas is cooled, and the greater part of the tar and water, which are the chief liquid products, are condensed and separated. It then passes first through a box filled with perforated trays containing slaked lime, and afterwards through a similar box containing ferric hydrate. The lime removes CO_2 and other acid impurities and the ferric oxide hydrogen sulphide, from which the sulphur can afterwards be removed. The gas is then carried to immense reservoirs called *gas-holders*, from which it is distributed for use.

The products of the distillation of coal may be roughly classified as follows. The list is not complete :

- | | | |
|---|---|--|
| 1. Coke (carbon). | Remains in the retort. | |
| 2. Tar and water. | Mostly removed when the gas is cooled. | |
| 3. Gases and vapours <i>injurious</i> to the gas. | Ammonia. | { Chiefly remains in the water in combination with the following : |
| | Carbonic anhydride. | |
| | Sulphurous " | { Partly remain in the water as ammonium salts. Partly absorbed by the lime. |
| | Sulphuretted " hydro- | |
| | gen. | |
| | Cyanogen. | |
| | Carbon disulphide. | Very difficult to separate. |
| 4. Diluents. | Hydrogen. | { Do not contribute to luminosity. |
| | Methane. | |
| 5. Useful. | Acetylene. | { Contribute luminosity. |
| | Ethylene. | |
| | Vapours of liquid and solid hydro-carbons | { Ditto. Also cause the smell of the gas. |

Flame.—What we ordinarily call a *flame* is the space in which gases are combining with oxygen. Flame is only produced by the union of *gases*. This at first sight appears startling. It is natural to ask whether the burning of a coal or a candle is not the burning of a *solid*, and the burning of spirit or oil, of a liquid. But, in reality, we do not burn these substances, but the gases or vapours which they give off when heated. Charcoal burns without flame, because it gives out no vapour. In burning phosphorus, sulphur, or spirits of wine, we really burn

the *vapours* of these substances, and in burning coal, wax, wood, oil, or paper, we burn the *gases* which are produced by a destructive distillation of those substances.

A candle is a miniature gas factory. In it wax or tallow is *distilled*, and the gas so obtained is burned. Blow out a candle and apply a light to the white smoke which ascends from the wick. It will ignite, and the flame will run down and settle on the wick.

It has been previously stated that hydrogen burns very easily, and with a flame, while carbon burns more difficultly, and without flame; this explains why fuel burns with a flame at the commencement of the combustion, but finally only glows; it is the gases which *first* burn with *flame*, and afterwards the carbon, with a mere glow, without flame. All combustible substances that contain hydrogen and carbon burn in a similar manner. Burning wood presents a simple illustration of this fact.

The flame of alcohol consists of two parts; the dark central part is alcohol vapour, and the bright envelope

Fig. 67. around this is alcohol vapour uniting chemically with the oxygen of the air. The tapering form of the flame is owing to the ascent of the hot gases, and the rushing in of cold air from below. The alcohol is drawn up from the lamp by the capillarity of the wick; it burns with little light, but if a twisted wire or some other solid body be introduced into it, it will then emit



light. If a thin wire is placed across the flame, it will be heated to redness near the margins of the flame, while in the interior it will remain dark; consequently, the external part is much hotter than the central part of the flame. The point of greatest heat is indicated by the mark in the figure, and vessels to be heated over the spirit-lamp should never be placed below this point. This may be rendered very evident by applying a lucifer-match to this part of the flame, when it will take fire at once; but not so quickly if thrust into the centre of the flame. The flame of hydrogen is similar to that of alcohol.

In the *flame of a candle or lamp*, three parts can be distinguished; in the *middle* (*a*, Fig. 68), the dark centre, consisting of illuminating gas (decomposed tallow);

around this (b), the luminous cone, consisting of burning hydrogen, intimately mixed with carbon or heavy hydrocarbons at a white heat; and on the very outside (c), Fig. 68. a thin, scarcely perceptible veil, in which carbon is burning. If we imagine the flame to be cut horizontally through the centre, it would present nearly the same appearance as in Fig. 69. The middle circle is illuminating gas, the hydrogen of which burns first, and the great heat thus evolved brings the carbon or hydrocarbons to a white heat (this is indicated by the second circle); and finally, in the exterior circle, the carbon is burnt. The heated matter in the second ring imparts to the flame its illuminating power, just as the glowing wire rendered the alcohol flame luminous. If a cold knife be introduced into the flame, a portion of the carbon will be so much cooled that it cannot burn, and will be deposited upon the knife in the form of soot. If a wire be held through the flame, the glowing part at the hot margins will remain clear, while soot will be deposited upon that part of it which is in the interior of the flame.

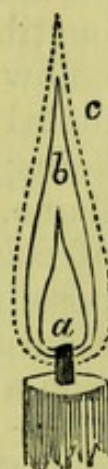
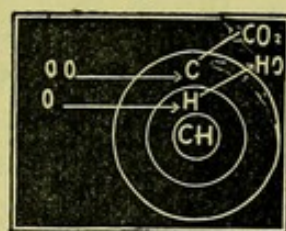


Fig. 69.



The *brightness* of a flame depends, either upon the presence of a solid body, which glows in the flame, or of dense hydrocarbons undergoing combustion in it; if these be only heated to redness, the flame will give out a smoky red light, but, on the contrary, a brilliant light when heated to a white heat.

The blowpipe.—In its simplest form, this valuable little instrument consists of a tapering metal tube bent at a right angle near its end, and terminating in a fine jet. With it a current of air can be forced through a small flame, sending it down horizontally, increasing its temperature very much and altering its character. The blowpipe flame consists of two cones: an inner one, which is blue, and has the power of removing oxygen from many metallic oxides, and an outer one, which is pale yellow, and has the opposite power of promoting oxidation. The inner cone, which is much hotter than the other, is called the *reducing flame*; the outer, the *oxidizing flame*. The

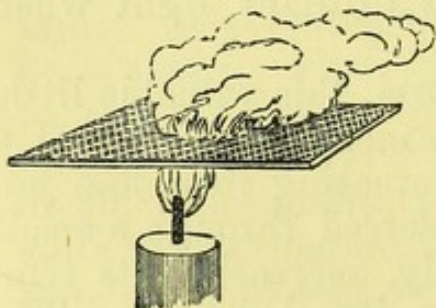
flame of the oxyhydrogen blowpipe (page 133) is similar to the ordinary blowpipe flame in structure.

Temperature necessary for flame.—A certain definite temperature is necessary for the combustion of each combustible substance, and if the temperature be reduced below that point the combustion ceases. A piece of cold metal will extinguish a candle flame, because the metal, being a good conductor, reduces the temperature of the flame. When we “blow out” a flame, the sudden current of cold air exerts a similar cooling action. It is impossible to blow out the flame of phosphorus or carbon disulphide, because those substances enter into combustion at a very low temperature.

Sir Humphry Davy made the important discovery that the *fire-damp* (methane) of coal-mines requires a very high temperature for its combustion. From this knowledge he was led to invent the “safety-lamp,” which has done so much to mitigate the horrors of coal-mining. It is simply an oil lamp in which the flame is entirely surrounded by fine wire gauze. If this lamp is burned in an atmosphere containing fire-damp, or even ordinary coal-gas, the gas will enter and burn inside the lamp, often extinguishing the oil flame, *but the flame cannot pass through the wire gauze*, and an explosion is therefore impossible, unless the lamp is out of order, or improperly used.

The wire gauze acts by cooling the flame below the point required for combustion. Its

Fig. 70.



effect can be shown by pressing a piece down on an ordinary gas or candle flame. The flame is, as it were, crushed down by the gauze, though the gas passes freely through, and can be burned on the other side.

Smoky flames.—When the supply of air to a flame is insufficient, it *smokes*; that is, a part of the carbon escapes oxidation. A chimney often obviates this; the heat produces an upward current in it, whereby more air is drawn in. In the *argand burner* the flame is ring-shaped, and air enters through the centre, as well as outside the flame.

SILICON.

Si = 28. Formula unknown.

Silicon is the most abundant solid element. It is always found in combination, and is only isolated with great difficulty. It enters into the composition of so many minerals that the greater part of the science of mineralogy is occupied with the history of the silicon compounds. Silicon, like carbon and boron, may be obtained in two, and possibly in three conditions: amorphous, graphitoid and the so-called adamantine silicon. In the free state the element is too expensive to be of any practical importance. Like carbon and boron, it is insoluble in all agents that do not change it, and non-volatile.

Silicic anhydride, SiO_2 , also called silica and silex, is the only known oxide of silicon. It occurs native and tolerably pure, as quartz, sand, flint, and almost perfectly pure as rock crystal, which is found beautifully crystallized in transparent six-sided prisms. Carnelian, agate, jasper, opal, calcedony, and some other well-known precious stones, consist likewise of silica; their colours are chiefly owing to the presence of minute quantities of metallic silicates.

Three varieties of silicic acid analogous to the three phosphoric acids are recognised. They must be prepared by indirect means, for silicic anhydride will not combine directly with water. The normal, or *ortho*-silicic acid, H_4SiO_4 , has not been obtained free from water.

All silicates are insoluble in water except certain alkaline silicates.

Soluble Glass.—*Experiment 1.*—Heat a common flint stone in a fire, and when it is red-hot quench it in cold water. The sudden cooling so influences and diminishes the internal cohesion, that now the flint can easily be reduced to a fine white powder. Boil in a clean iron ladle two drachms of the powder with four drachms of caustic potash (potassium hydrate) and two ounces of water, for some hours, supplying fresh water occasionally as the other evaporates; then let the mixture stand in a corked bottle to settle. The silica dissolves in the potash solution

and forms with it a thick fluid, *potassium silicate*. Or one part of the powdered flint may be fused with four parts of dry sodium carbonate (common washing soda), in a crucible over the fire. The soda soon commences to effervesce, from the escape of carbonic anhydride, which is replaced by the silicic anhydride. The mass on being boiled with water and filtered furnishes a solution of *sodium silicate*. These soluble silicates are often called *soluble glass*, and their solutions, *liquor silicum*.

Liquor silicum has of late years been largely employed for the hardening of soft building stone, and the artificial preparation of hard stone from sand.

Experiment 2.—Add some hydrochloric acid to a strong solution of sodium or potassium silicate. A white gelatinous mass of silicic acid (probably H_4SiO_4) is thrown down. If this jelly is collected on a filter, washed well with water and dried, a white powder, consisting of pure silica in the amorphous state, is obtained.

Experiment 3.—If, on the contrary, the silicate solution be diluted with ten or twelve times its bulk of water, and then the alkali be removed with hydrochloric acid, no precipitate is produced, the silicic acid remaining in solution. This solubility is but a temporary state, as after the lapse of a certain time, depending on the proportion of silica present, the solution suddenly undergoes a remarkable conversion into a transparent jelly-like mass, the silicic acid passing into the insoluble modification. By the addition of more alkali it may be again re-dissolved and the experiment repeated.

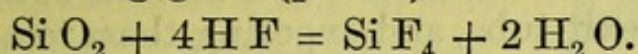
Almost all our springs, as well as our plants, contain small quantities of soluble silicates. If we evaporate spring-water, we find silica in the insoluble residue; and if we burn a plant, we obtain it in the ashes. Grasses, and the different kinds of grain, are particularly rich in silica, and for this reason they have been called *silicious plants*. Silica is to these plants what bones are to men,—the substance to which the stalk owes its firmness and stiffness. If the soil is deficient in soluble silica (or if there is not enough potassa, which renders the silica soluble), these properties will be wanting in the stalk, and it will bend. The horse-tail plant (*Equisetum*)

contains so much silica that it may be used for polishing wood. Silica is found even in the animal kingdom, particularly in the class of Infusoria, which are only visible under the microscope; the shells of many Infusoria are composed of silicic anhydride.

In its natural state, silica is so hard as to produce sparks with steel, and is quite insoluble in water and acids, except hydrofluoric acid.

Silicates.—The silicates are the most numerous and complicated of all salts. Their study, however, belongs to the science of mineralogy. Among important silicates may be mentioned felspar, mica, hornblende, serpentine and meerschauum. The different varieties of *clay* are silicates of aluminium. *Glass* is a mixture of various silicates, with excess of silica.

Silicon Fluoride, SiF_4 .—This interesting gas is obtained by the action of hydrofluoric acid on silica, or any silicate. To its formation the power which hydrofluoric acid possesses of etching glass (p. 122) is due:



It is immediately decomposed by water, with formation of silica and an acid called hydrofluosilicic acid, 2HF , SiF_4 .

Experiment 4.—Heat powdered glass, fluor-spar and sulphuric acid in a flask, and carry the gas produced to the bottom of a glass of water, in which some mercury has been placed. As the gas rises from the mercury it is decomposed by the water, and beautiful tubes of silica are built up. If it were not for the mercury, the mouth of the tube would soon be choked up by the silica, and the flask would then be apt to burst.

PART III.

METALS.



CHAPTER I.

METALLIC MONADS.

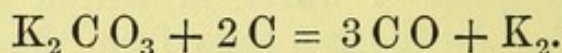
(Potassium, Sodium (Ammonium), Silver.)

POTASSIUM.

K = 39.

THIS interesting metal was discovered by Davy, who obtained it by decomposing caustic potash (potassium hydrate) by means of a powerful galvanic battery.

Potassium is now prepared by putting potassium carbonate and finely divided carbon into an iron bottle, provided with an iron exit-tube, and exposing them to a white heat. At this high temperature the carbon combines with the oxygen of the carbonate, forming carbonic oxide gas, which escapes. The liberated potassium is also converted into a green vapour, which is conducted into naphtha, where it condenses into a solid mass.



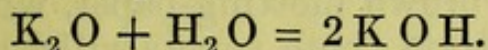
It has been shown, under Carbonic acid, that potassium, at a moderate heat, can withdraw the oxygen from the carbon; while here, at a higher temperature, a contrary action takes place. Similar incongruities in chemical actions are not unfrequent; they show that the affinities of bodies for each other are greatly altered by the temperature.

Experiment.—If a piece of potassium is cut with a knife, it presents a blue metallic lustre; but it immediately tarnishes on exposure to air, and soon becomes converted into solid *potassium oxide*. This experiment shows the softness of the metal, and also its great affinity for oxygen.

Consequently, to preserve potassium from oxidation, it is necessary to keep it in mineral naphtha, or some other

liquid which contains no oxygen. Its symbol, K, is derived from the Latin name *Kalium*.

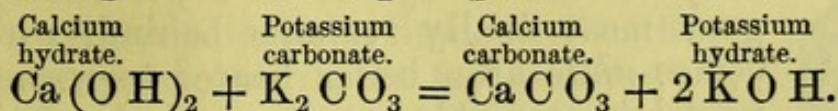
Potassium oxide, K_2O .—A small piece of potassium, when exposed to dry air, is converted into a white mass of potassium oxide. This substance combines eagerly with water, and forms potassium hydrate:



Potassium hydrate or *hydroxide*, KOH .—*Caustic potash*—*Potassa*.—This important substance is obtained when potassium acts on water (p. 105), but is more economically prepared as follows:—

Experiment 1.—Place half ounce of quick lime in a plate, sprinkle it with warm water, and let it stand until it is slaked, that is, until it becomes a fine dusty powder. Then put half an ounce of potassium carbonate into an iron basin with six ounces of water, boil it, and, during the boiling, gradually add half of the slaked lime by teaspoonfuls, stirring it at the same time. After the mixture has boiled for some time, put a teaspoonful of it upon a paper filter, and pour the filtrate into vinegar. If it effervesces, still more lime must be added; but if no effervescence ensues, pour the whole into a bottle, close it up, and let it remain quiet for some hours, that the sediment may subside. Decant the clear liquor, and preserve it in a well-stoppered bottle. It is a solution of *potassium hydrate*.

Slaked lime is calcium hydrate: when it is boiled with potassium carbonate, calcium carbonate and potassium hydrate are formed by double decomposition, and the former, being insoluble, is precipitated:



Experiment 2.—Evaporate a portion of the potash solution in a clean iron vessel (glass and porcelain are attacked by it); all the water is expelled and a white mass finally remains behind, dry potassium hydrate. This may be melted at a stronger heat, and cast into sticks or plates.

Fig. 71.

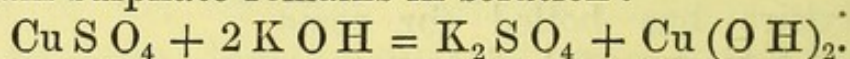


Experiment 3.—Expose some dry potash to the air; it will soon become moist; it will deliquesce, and on longer exposure will effervesce upon the addition of an acid. Potash absorbs both water and carbonic anhydride from the air, and is then converted into potassium carbonate.

Experiment 4.—Heat in one test-tube some white, and in another some brown blotting-paper, with some solution of potassium hydrate; both papers will be decomposed and dissolved, the vegetable fibres of the white paper (linen or cotton) more slowly than the animal fibres of the brown paper (wool). Potash exerts a very *corrosive* action, especially on animal substances. The slippery feeling caused by rubbing lye between the fingers is owing to a gradual solution of the skin.

Experiment 5.—Boil in a test-tube a little tallow or fat with a solution of potassium hydrate; a union gradually takes place; soap is formed. The soap prepared from potash remains soft, and is called soft-soap.

Experiment 6.—Dissolve a piece of blue vitriol (copper sulphate) in water, and add to it some solution of potash. The copper is precipitated as a blue hydrate, and potassium sulphate remains in solution:



Potassium hydrate being a strong base precipitates many other metals from their salts as hydrates, by an interchange of the respective radicals.

Potassium Carbonate, K₂CO₃.

Experiment 1.—Fit into a funnel a filter of blotting-paper, and place upon it a handful of *wood ashes*, and gradually pour hot water over them. The liquid filtered through has an alkaline taste, and turns red test-paper blue. If you evaporate it to dryness in a porcelain dish, a grey mass finally remains behind, which becomes *white* after being heated to redness in a porcelain crucible; it is *pearlash*, or impure potassium carbonate. In those countries where wood is abundant—in America, Russia, &c.—it is prepared in a similar manner on a large scale, and is an article of great demand in commerce.

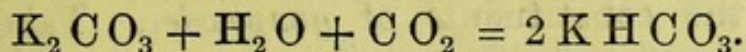


It can be purified by recrystallization.

Experiment 2.—Put one portion of potassium carbonate in a vessel, and let it stand in a dry apartment, and put another portion in a cellar. The former becomes moist; the latter deliquesces. Both attract water from the air, but that in the dry atmosphere of the room less than that in the damp air of the cellar. Potassium carbonate is a very *deliquescent salt*.

Experiment 3.—Boil for some time, in a vessel containing a quarter of an ounce of potassium carbonate and two ounces of water, a piece of grey linen, and some dirty or greasy linen or cotton rags; the liquid will become of a dark colour, while the rags are made white and clean. Dirt, as it is commonly called, is dust, which adheres to the skin, garments, &c., particularly after they have become moistened by perspiration, or have come in contact with greasy or other adhesive substances. These last-mentioned substances may be dissolved and removed by pearlash; on this depends the various application of this substance in cleaning and washing.

Hydrogen potassium carbonate, HKCO_3 .—This salt, which is commonly termed bicarbonate of potash, differs from the preceding compound in having but one atom of the hydrogen in carbonic acid replaced by potassium. It consequently belongs to the class called acid salts, and is prepared by passing a current of carbonic anhydride through a solution of ordinary potassium carbonate until no more is absorbed:



Potassium sulphate, K_2SO_4 .—Dissolve half-an-ounce of potassium carbonate in two ounces of warm water, and then add diluted sulphuric acid in very small quantities until no more effervescence is produced on stirring.

Filter and evaporate the liquid until a film of crystals appears on the surface, then set it aside. The hard crystals obtained (six-sided double pyramids) are potassium sulphate; they are sparingly soluble in water, and have a somewhat bitter taste.

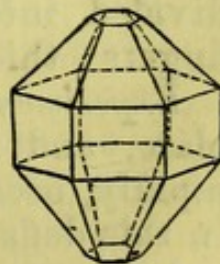


Fig. 73.

Hydrogen, potassium sulphate, HKS_4 , sometimes called

bisulphate of potash, is obtained as a secondary product in the preparation of nitric acid from saltpetre. It is an extremely acid salt, and much more soluble than the neutral sulphate.

Potassium Nitrate, KNO₃.

Experiment 1.—Dissolve half an ounce of potassium carbonate in one ounce of hot water, and neutralise with nitric acid; afterwards boil and filter the liquid, and set it aside to cool; prismatic crystals of potassium nitrate will be deposited from it, which have a cooling taste, and undergo no alteration in the air.

This salt is known as *nitre* and *saltpetre*; after fusion, it solidifies to a crystalline mass, sometimes called *sal-prunella*.

Experiment 2.—Throw a little nitre on a red-hot coal, it will cause a brisk sparkling; it *deflagrates*. The nitre is decomposed and gives up its oxygen to the coal, and thus causes it to burn more energetically. The hard saline mass, congealed from its melted state, remaining on the coal, has a basic reaction, and effervesces with acids; it is potassium carbonate. In order to render substances more inflammable, they are often soaked in a solution of nitre; as for example, tinder, touch-paper, &c.

Experiment 3.—Mix thoroughly in a mortar fifteen drachms of powdered nitre, three drachms of charcoal-powder, and two drachms of sulphur; this mixture contains the constituents of *gunpowder*. Take a little on the point of a knife, put it on a stone, and ignite it with a match; a brisk deflagration will ensue. Knead the rest of the powder, with some drops of water, into a paste, and squeeze it through a tin colander or a sieve. The thread-like mass thus obtained is, when partly dry, divided into small grains by gently rubbing with the fingers; this is gunpowder.

Experiment 4.—Place some gunpowder upon an iron plate, and ignite it; the deflagration takes place more rapidly, because the materials are more intimately mixed. In this deflagration there are evolved from the charcoal and the nitre *carbonic anhydride* and *nitrogen*, two gases

which instantly occupy a much greater space than before. Sulphur not only causes an easier ignition of the gunpowder, but it causes a stronger evolution of gas; since it combines with the potassium of the nitre, forming potassium sulphide, and liberates carbonic anhydride. The change which occurs is somewhat complex.

If the deflagration of the gunpowder takes place in a confined space, as in a gun-barrel, the explosive violence with which the two gases are suddenly expanded is strong enough either to project the ball or to burst the gun. The potassium sulphide remaining soon becomes moist in the air, and then emits the odour of sulphuretted hydrogen; at the same time, the iron is blackened by the formation of sulphide of iron upon the surface.

If nitre be heated with sulphuric acid, nitric acid escapes.

Animal substances are preserved from putrefying by nitre; it is therefore used in salting meat.

The manufacture of nitre was formerly conducted in a very peculiar manner. Animal substances, for instance, pieces of flesh, hides, hair, &c., are mixed with wood-ashes and lime, and then moistened with water or urine, and suffered to putrefy slowly. Animal substances are rich in nitrogen, which, during putrefaction, is set free in the form of ammonia (NH_3); this, after a time, unites with the oxygen of the air, forming nitric acid, which acid is immediately neutralised by the potassium of the wood-ashes. This process is called *nitrification*. If animal substances decay without the presence of potash, or some other strong base, no nitric acid, but only ammonia, will be produced.

After the completion of the putrefaction, water is added to extract the nitre, and the solution being evaporated it is deposited in crystals. Nitre beds, so called, are prepared in this way. We formerly obtained nitre from the East Indies, where it is found as an incrustation on the surface of the soil.

Potassium chlorate, KClO_3 .—This salt is very similar to nitre, but it is more easily decomposed. By mere heat it is resolved into oxygen and potassium chloride; therefore it is used for the preparation of oxygen. Its preparation has already been described (p. 138).

Experiment 1.—When thrown on glowing coals, it deflagrates still more briskly than nitre; the oxygen, as it is liberated, occasions a very energetic combustion of the coal. This salt cannot be employed in the preparation of gunpowder, as the rapidity with which it explodes would be too much for the guns; yet, on this very account, it is extremely serviceable in fireworks, especially for producing variegated fires. The *greatest caution* must be observed in pulverising and mixing it with combustible materials, as it may explode by merely rubbing or pounding it. The mixing of it with other substances must always be done with the fingers.

Experiment 2.—Introduce a crystal of potassium chlorate into a beaker-glass, and add a few drops of alcohol, and afterwards a drop of sulphuric acid; the sulphuric acid expels chloric acid, which is immediately decomposed, and there is so great an evolution of heat as to set fire to the alcohol.

Experiment 3.—Mix a very little potassium chlorate between the fingers with about half as much flowers of sulphur and throw the mixture into sulphuric acid, contained in a beaker-glass; a brisk crackling ensues, and the sulphur takes fire.

Potassium Chloride, KCl.

This salt is found in very large quantity at Stassfurth in Germany and is now the most important source of potassium compounds.

Experiment 1.—Dissolve half an ounce of potassium carbonate in water, and neutralise with hydrochloric acid; upon concentrating the solution, cubic crystals will be obtained, having a taste similar to that of common salt.

Potassium Iodide, KI.

This salt likewise crystallizes in cubes, is easily soluble in water, and is employed in medicine as a valuable remedy.

Experiment 1.—To prove that iodine is really contained in this white salt, heat a small portion of it in a test-tube with a little black oxide of manganese and some drops of sulphuric acid, when violet fumes will be evolved. If

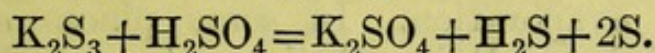
common salt is treated in the same manner, chlorine, as is known, will be given off. The chemical action is the same in both cases.

Potassium Tersulphide, K₂S₃.

Sulphur combines with potassium in five different proportions, forming ill-defined compounds. One of the most important of them is prepared as follows:

Experiment 1.—Put a mixture of one drachm of sulphur and two drachms of dry potassium carbonate into an iron ladle, cover it with a strip of sheet iron, and heat it until the effervescence has ceased and the mass flows quietly. The fused mass has the colour of liver, and on this account has received the name *liver of sulphur*: pour it upon a stone, and if it should inflame, cover it with a vessel to extinguish it. On exposure for some time to the air it becomes greenish and moist, and evolves an odour like that of rotten eggs. It consists chiefly of potassium tersulphide. A similar preparation has already been described (p. 141).

Experiment 2.—Pour water into a test-tube containing some liver of sulphur; you obtain a yellowish-green solution. If to this you add diluted sulphuric acid, a strong evolution of *hydrogen sulphide* takes place, and the liquid becomes milky from the precipitation of two-thirds of the sulphur (milk of sulphur):



The same thing is effected, though far more slowly, by the carbonic anhydride of the air, and thus is explained why the liver of sulphur (as well as the residue left on the combustion of gunpowder) emits a smell like that of rotten eggs when it is left exposed to the air.

The other sulphides of potassium can be formed by varying the proportions of the potassium carbonate and sulphur.

SODIUM.

Na = 23.

Sodium was discovered by Davy soon after the isolation of potassium. Like potassium, it is obtained by reducing

its carbonate with carbon. It is much easier to prepare than potassium.

Sodium is a soft, white, and easily fusible metal, lighter than water, and very similar to potassium in chemical characters.

Experiment 1.—The action of sodium on water has already been described (p. 104). It is closely analogous to that of potassium.

Sodium oxide, Na_2O , and *Sodium hydrate*, NaOH , are obtained in the same way as the potassium compounds, and are so analogous to them that the descriptions given of the latter are applicable, with few exceptions to either.

Sodium chloride, or *common salt*, NaCl .—Salt has twice previously been artificially prepared; namely, once from sodium and chlorine (page 113), and again from soda and hydrochloric acid (page 117); its constituents are accordingly already known. It has the formula NaCl .

The earth and sea abound in common salt; it may therefore be easily procured in large quantities. In many places it is found in the interior of the earth, in immense beds, from which it is broken up and dug out. This salt looks like a transparent stone, and is, therefore, called *rock-salt*. In those places where the rock-salt is mixed with stones and earth, a hole is bored in the middle of the bed, and water is let into it. The water is pumped out again as soon as it has become saturated with the salt, and is again expelled by evaporation. In some places springs are found containing salt in solution—the so-called *natural brine springs*. These are always occasioned by the water permeating the earth over a bed of rock-salt, and appearing as a spring at some lower level.

In hot countries, salt is also prepared from sea-water, which is evaporated in shallow tanks by the heat of the sun. It is called *bay-salt*, and has a bitterish taste, owing to the presence of salts of magnesium. Sea-water contains rather more than 3 per cent of common salt.

Salt is indispensable to the life of animals, and is therefore a constant and necessary ingredient of food. It is also used for preserving animal and vegetable substances, because it has the power of preventing putrefaction or decay. Meat and fish are salted down, and wood for the

purpose of building is rendered more durable by being impregnated with salt. Such substances are called *antiseptics*.

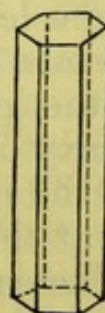
Experiment 1.—Dissolve one ounce of salt in two ounces and three-fourths of cold water; the water will dissolve no more. Repeat the experiment, using hot instead of cold water; the result is almost exactly the same. Common salt has the remarkable property of *being nearly as soluble in cold as it is in hot water*. A larger quantity of almost all other salts is dissolved by hot than by cold water. Put one of these solutions in a warm place; by the gradual evaporation, regular transparent cubical crystals of common salt are formed. Boil down the other solution, quickly stirring it all the while; it yields a granular, opaque, saline powder. Salt is prepared as last described on a large scale, and hence the granular state of common salt.

Experiment 2.—If you expose a solution of salt in an open place during the extreme cold of winter (-10°C. or $+14^{\circ}\text{F.}$), transparent prismatic crystals will be formed, which contain more than one-third of *water of crystallization*, $\text{NaCl}, 2\text{H}_2\text{O}$. When placed on the hand they quickly become opaque and deliquesce into a syrupy mass, in which numerous small cubic crystals may be perceived.

Experiment 3.—Heat some common salt on platinum foil or a knife, it will crackle briskly, and part of it will be thrown off the foil; that which remains *melts* when the foil becomes red hot. The crackling (decrepitation) proceeds from a trace of water, which has remained in the interstices of the crystals; on being heated it expands and bursts the crystals asunder.

Sodium sulphate, Na_2SO_4 .—As most of the potassium salts and potassium are prepared from potassium chloride, or carbonate, so most of the sodium salts are prepared from common salt. In the latter case, however, an indirect process must often be resorted to, since chlorine is not so easily removed from sodium as carbonic acid is from potassium. The sodium chloride must first be converted into sodium sulphate. We are already acquainted with this salt, as it remains in the retort after the preparation of hydrochloric acid,

Fig. 74.



when common salt is heated with sulphuric acid. It was formerly taken as a popular medicine, under the name of Glauber's salts, so called from its discoverer, the physician Glauber. We find it also in many mineral waters; for instance, in the Cheltenham, Püllna and Hunjadi Janos waters. It is readily soluble, crystallizes in prisms, and has a saline, bitter taste.

Experiment 1.—Place half an ounce of crystals of sodium sulphate in a warm place; they soon become covered with an opaque white coating; they *effloresce*. The powder obtained weighs hardly a quarter of an ounce. That which was lost was water. Crystals of sodium sulphate contain more than half their weight of water of crystallization. Their composition is expressed by the formula $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$. The transparency of the salt is lost when chemically combined water is separated, but reappears when the anhydrous salt is dissolved in water and re-crystallized. Salts which effloresce must be kept well-corked up.

Experiment 2.—If a crystal of sodium sulphate is heated on charcoal before the blowpipe, it soon melts, because it dissolves in its water of crystallization (watery fusion); it becomes dry as soon as the water is expelled; but finally it melts for the second time when heated to redness (igneous fusion). Those salts which contain no water of crystallization can undergo only the latter kind of fusion.

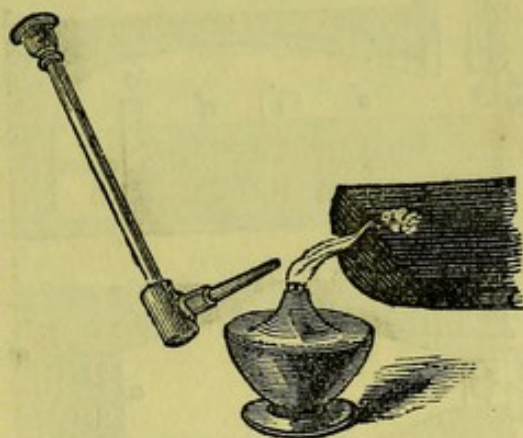
Experiment 3.—Heat in a small flask half an ounce of water to 91°F. (33°C.), and keep it at this temperature, gradually adding crystallized sodium sulphate, as long as they are dissolved, amounting to about an ounce and a half. If a stronger heat be now applied to the saturated solution, a salt will separate (anhydrous crystals, Na_2SO_4); if you let it cool, a salt will likewise separate (hydrated crystals, $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$);—furnishing another example of the great influence exerted by temperature on the affinity of water for other substances. Sodium sulphate has the peculiar property of being most soluble in water, not at the boiling point, but at a lower temperature.

The curious manner in which a hot saturated solution of the salt can be cooled without immediate crystallization taking place, has already been noticed (p. 10, Exp. 3).

Experiment 4.—If you dissolve crystallized sodium sulphate in water, *cold* is produced; but if, on the contrary, you dissolve the anhydrous salt in water, then *heat* is produced. You will observe exactly the same phenomena if you perform this experiment with sodium carbonate, taking first the crystallized and then the calcined salt. Whence the source of this heat? It is due to a part of the water combining with the anhydrous sulphate, or carbonate, as water of crystallization.

Sodium sulphide, Na₂S.—*Experiment 1.*—Mix a small portion of anhydrous sodium sulphate with a little charcoal powder, and heat the mixture

Fig. 75.



on charcoal before the blow-pipe; they will melt with brisk effervescence into a brown mass, which dissolves in water, forming a yellowish liquid. The charcoal, when heated to redness, abstracts the oxygen, and forms with it carbonic oxide gas, which escapes with effervescence; sodium and sulphur remain behind, combined with each other. That is, the charcoal deoxidizes the sodium sulphate, or reduces it to sodium sulphide.

If you drop hydrochloric or diluted sulphuric acid into the solution, the disagreeable smell of sulphuretted hydrogen will be given off.

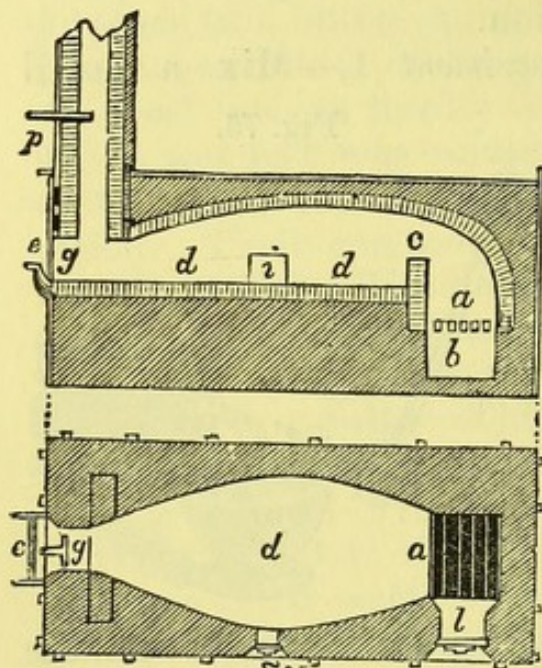
Sodium Carbonate, Carbonate of soda, Na₂CO₃.

Experiment 1.—Prepare some more sodium sulphide in the manner just described, rub it in a mortar with the adhering particles of charcoal and with about its own weight of chalk, and ignite it again before the blow-pipe. Warm the baked saline mass in water, and then filter the liquid. A grey powder remains behind, which, when drenched with hydrochloric acid, evolves sulphuretted hydrogen; it is calcium sulphide. The liquid, after being evaporated on a shallow glass dish, leaves behind a white powder, which has an alkaline reaction and effervesces

with hydrochloric acid, but yet without emitting any disagreeable odour; it is *sodium carbonate*. The sulphur has thus passed to the calcium of the chalk, while the carbonic acid radicle of the chalk has passed to the sodium.

As sodium carbonate possesses almost exactly the same properties as potassium carbonate, and can be advantageously employed instead

Fig. 76.



of the latter in washing and bleaching, and also in the manufacture of glass and soap, it is now manufactured on an enormously large scale in chemical works. The process pursued is essentially the same as that already described, except that the two operations, described separately above, are united into one; the chalk or limestone is added, in the first place to the sodium sulphate and charcoal, and the whole mass is heated. This is done in large oven-shaped furnaces, represented in the figure; *a* is the grate, *b* the ash-pit, *p* the chimney, *d d* the hearth for receiving the mixture, *i* the aperture for throwing in the mixture, and *g* an opening for stirring it and scooping it out. They are called *reverberatory-furnaces*, because the heating is effected, not by the fuel itself, but by the flame, which is reverberated after passing over the bridge *c*; they possess this important advantage, that the ashes of the coal or peat do not become mixed with the substance to be heated. In some countries sodium carbonate is also obtained from the ashes of marine plants (kelp).

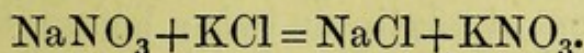
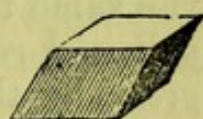
Sodium carbonate occurs in commerce, either *crystallized*—it then contains more than half its weight of water of crystallization. $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$, and effloresces very readily—or *calcined*, consequently anhydrous. The latter, accordingly, when it occurs pure, is of more than twice the strength of the crystallized. Sodium carbonate is

easily soluble in water. Many mineral waters contain great quantities of it in solution.

Hydrogen sodium carbonate, HNaCO_3 , is chiefly used for preparing effervescing powders. It is prepared like the similar potassium salt (p. 199), and is commonly called bicarbonate of soda.

Sodium nitrate, NaNO_3 .—Dissolve some sodium carbonate in water, neutralize with nitric acid, evaporate the solution, when crystals will separate, having the form of an oblique rhombic prism; they are nitrate of soda. They deflagrate on charcoal like potassium nitrate, and have the greatest similarity to it in other respects. Extensive beds of this salt are found in America, whence large quantities are exported, under the name of *Chili salt-petre*; and it is substituted for the more costly nitre in the manufacture of nitric acid and some of its salts. But it does not answer for making gunpowder, as it is deliquescent. Saltpetre is prepared from it by the addition of the Stassfurth potassium chloride and subsequent crystallization.

Fig. 77.



Sodium Phosphate, Na_2HPO_4 .

Experiment.—Neutralise sodium carbonate, dissolved in water, with phosphoric acid; filter the liquid, and evaporate the filtrate until a film forms on the surface; on cooling, transparent crystals will be deposited, which contain more than half their weight of water of crystallization: $\text{Na}_2\text{HPO}_4 \cdot 12\text{H}_2\text{O}$. They easily effloresce, and yield a *yellow* precipitate with a solution of silver nitrate.

Other phosphates of sodium of less importance are known.

Sodium di-borate, or *Borax*, $\text{Na}_2\text{B}_4\text{O}_7$.—The hard colourless crystals commonly called borax are an irregular sodium salt of boric acid (p. 174). Tincal is a native borax found in China and Thibet. Borax contain ten molecules of water of crystallization: $\text{Na}_2\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}$.

Experiment 1.—Heat some powdered borax upon a small loop of platinum wire before the blow-pipe; it will puff up and swell in its water of crystallization, and be con-

verted into a porous spongy mass; on being further heated, it fuses to a transparent *bead*. Moisten this bead with the tongue, apply it to litharge so that some of the latter may adhere to it, and again hold it in the outer flame of the blow-pipe; the litharge is dissolved; the bead remains colourless and transparent. If you now substitute for the litharge other metallic oxides, you will likewise observe that the oxides will dissolve, but that at the same time the bead will be coloured by them; namely, *yellowish-red*, by sesquioxide of iron and oxide of antimony; *green*, by the oxide of chromium; *blue*, by oxides of copper and cobalt; *violet*, by a small portion of oxide of manganese; and *brownish-black*, by an excess of manganese.

On account of this property which borax has of dissolving metallic oxides, it is used in chemistry as a blow-pipe test for the detection of metallic oxides, and in the trades for *soldering*, or joining one metal with another.

Sodium Silicates (Glass, &c.).

Experiment.—Melt some potassium or sodium carbonate in a small loop of platinum wire before the blow-pipe, and then add a little finely pulverised sand; upon placing it again in the blow-pipe flame, effervescence will ensue, and afterwards a clear bead will be formed. If the proportion of sand used be small, the glass formed will dissolve in water on long-continued boiling; it is then called *soluble glass* (p. 194). If more sand is taken, a glass is obtained which is very difficult to dissolve in water. To make a glass which shall be entirely insoluble, not only in water but also in acids, beside alkalies, some other earth or metallic base—for instance, lime or litharge—must be added. Common glass is thus manufactured in glass-houses:

The materials which are chiefly employed in the manufacture of glass are quartz, flint, or sand, potassium carbonate or wood-ashes, sodium carbonate, lime or chalk and litharge or lead oxide. These substances, after being pulverised, are mixed together, thrown into earthen pots, and heated in a furnace until the mass forms a uniform fluid. In this state it may be moulded like

wax, cut and bent, pressed into moulds, and blown, and may accordingly be manufactured into all possible shapes and forms; on cooling, it becomes hard and brittle. In order to diminish in a measure the brittleness, the glass must be cooled very slowly (*annealed*). Glass vessels that are rapidly cooled often crack when they are carried from a warm into a cold room; this defect may, to a certain degree, be corrected, by gradually heating the vessels in water till it boils, and then allowing it to cool very slowly.

For colouring and painting glass various metallic oxides are employed. The milk-white colour which we observe in the opaque glass of the lamp-screens, and in the enamel of the dial-plate of watches, is produced by finely-ground bone-earth or meta-stannic acid—neither of which substances is dissolved by the vitreous mass, but only mixes with it mechanically, and renders it opaque, as chalk does water. Glass is ground by sand and emery, polished by ferric oxide and tripoli, etched by hydrofluoric acid, and very easily perforated by the point of a three-cornered file, which should be frequently moistened with oil of turpentine.

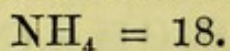
The two principal kinds of glass are—

Crown or Bohemian glass, consisting of sodium and calcium silicates.

Flint or crystal glass, consisting of potassium and lead silicates.

Common *bottle-glass* contains the same ingredients as crown glass, with the addition of ferric oxide, which imparts to it a brownish-yellow colour; of ferrous oxide, which gives it a green tinge; of alumina, &c. The iron is contained in the impure materials (yellow sand and wood-ashes) used in the preparation of the ordinary sorts of glass.

AMMONIUM.



Although ammonium is a compound radical and not an element, its compounds, which are often called *ammonia-salts*, are so similar to those of potassium and sodium, that they are generally placed among the metals. Their constitution has already been explained (p. 158).

Ammonium hydrate, NH_4OH .—The solution of ammonia gas in water has been supposed to contain this compound. The solution is strongly alkaline and caustic, like that of potassium, or sodium hydrate.

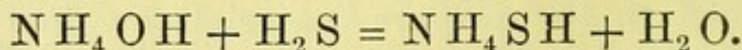
Ammonium chloride, NH_4Cl . *Sal-ammoniac*.—Formed by the direct union of ammonia and hydrochloric acid (p. 158). It is generally prepared from the “ammoniacal liquor” of gas-works (p. 157). It is a tough, fibrous solid, which may be purified either by crystallization from water or by *sublimation*.

Experiment 1.—Heat some impure sal-ammoniac in a dry test-tube. It will volatilize, and the colourless vapour will condense on the sides of the tube as a white mass, which however, is still apt to contain iron.

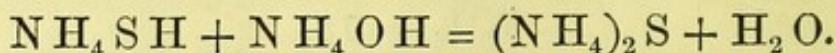
Experiment 2.—Dissolve some crude sal-ammoniac in water, and add a few drops of *ammonium sulphide*. If iron be present, a black precipitate will appear, and will subside after a time. The filtered liquid may then be evaporated to dryness, the mass dissolved in water, the solution filtered and evaporated until it crystallizes. The crystals will consist of pure ammonium chloride.

Ammonium sulphide, $(\text{NH}_4)_2\text{S}$, and *Sulphydrate*, NH_4SH .—These compounds correspond to the hypothetical oxide and hydrate of ammonium, $(\text{NH}_4)_2\text{O}$, and NH_4OH .

Experiment 1.—Pass sulphuretted hydrogen gas, prepared as described on page 143, first through a bottle containing a little water to remove impurities, and then into some dilute ammonia. The gas will be absorbed in large quantity. When the ammonia is saturated, there will remain a solution of *ammonium sulphydrate*:



If this solution is mixed with a quantity of ammonia equal to that first employed, the *sulphydrate* becomes *sulphide*:



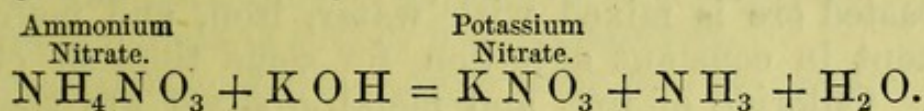
This sulphide precipitates many metals, and is therefore a valuable test. When freshly prepared it is colourless, but soon turns yellow from the formation of a higher sulphide. The change does not often hinder its usefulness as a test.

Ammonium carbonate.—*Experiments.*—Mix an ounce of chalk (calcium carbonate) and half an ounce of ammonium chloride in powder. Heat the mixture in a dry Florence flask with the neck broken off, over which is inverted a small beaker. Double decomposition takes place; calcium chloride remains in the flask, and the volatile *ammonium carbonate* sublimes and condenses in the beaker.

This is the well-known *smelling salts*. It has, as its common name implies, a strongly ammoniacal odour. In reality the salt has a complex constitution, but the student may safely regard it as *ammonium carbonate* $(\text{NH}_4)_2\text{CO}_3$.

Ammonium nitrate, NH_4NO_3 , has already been described (p. 158).

All ammonium salts, when warmed with solution of potash or milk of lime, give off ammonia gas. They may readily be identified by this means:



SILVER.

$\text{Ag} = 108$.

Silver, though a monad, has but little analogy with the others metals of this group. It sometimes occurs native (uncombined), but more frequently as sulphide, either pure or associated with the sulphide of lead or copper. It is also found as chloride (horn-silver). Several processes are adopted for its extraction. From galena (lead sulphide) it is obtained by reducing both metals to the metallic state by roasting and smelting the ore with charcoal. The silver is then separated by a process called *cupellation*. The alloy is placed on a hearth or *cupel*, made of bone ash, and strongly heated in a current of air. The lead is rapidly converted into oxide, which, fusing, is partly volatilized, and partly absorbed by the porous hearth, leaving the silver in the metallic state. If the lead be poor in silver, the alloy is melted, and allowed to cool slowly in iron pots. The portions which solidify first consist chiefly of lead, and being removed with a perforated ladle, the remaining lead becomes gradually much richer in silver. Complete separation is finally effected by

cupellation. This method, known as Pattinson's process, can be used economically, even when the lead contains less than $\frac{1}{5000}$ its weight of silver. Zinc is sometimes used instead of lead. Silver is obtained from the copper ores by reducing them to the metallic condition. The copper containing the silver is fused with lead, and the alloy, after solidifying, is gradually heated. The lead and silver melt first, and run off, leaving the copper. The silver is finally isolated by cupellation.

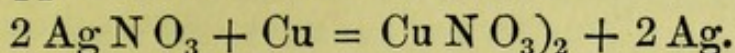
Silver is often extracted by means of mercury from the ores containing pure silver or silver sulphide, but no lead. But in the case of silver-glance (silver-sulphide) the metallic silver must first be separated from the sulphur. This is done by two operations. In the first, the stamped ore is roasted with common salt, by which process silver chloride and sodium sulphate are formed; in the second, the roasted ore is mixed with water, iron, and mercury, and kept in constant agitation for some time in closed casks. Chloride of iron and metallic silver are thereby formed, the latter of which is dissolved in the mercury. The excess of mercury is then filtered off, and a solid silver amalgam is obtained by subjecting it to pressure, and the mercury is at last completely removed from the amalgam by distillation. Several other methods for the separation of silver are, or have been used.

As pure silver is very soft, and would quickly wear out in using, it is generally alloyed with copper, whereby it is rendered harder, without losing its ductility. If the proportion of copper is only one-fourth, the silver still retains its beautiful white colour; but if more copper is added, the alloy becomes yellow, and finally red, by use. The standard silver, used in England for coinage, consists of $9\frac{1}{4}$ parts of pure silver and $\frac{3}{4}$ of copper.

Silver nitrate, Ag NO_3 .—*Experiment 1.*—Heat a sixpence in some dilute nitric acid until it is dissolved. Silver nitrate is formed, but is contaminated with copper, as the blue colour of the solution betokens.

Experiment 2.—Place a bright sheet of copper in the silver solution, and let it stand for a few hours. The copper immediately becomes covered with a black powder, which soon increases to a mass of brilliant metallic crystals

of silver. The copper decomposes the silver nitrate forming copper nitrate and free silver :



By washing the crystals well with water in a filter, and drying them, pure silver is obtained, which, by being again dissolved in nitric acid, and evaporated to dryness at a gentle heat, furnish a white mass of pure silver nitrate. This may be dissolved in water, and crystallized, or fused and formed into sticks, when it constitutes *lunar caustic*. The crystals are anhydrous.

Experiment 3.—Place a small piece of silver nitrate upon charcoal, and heat it before the blow-pipe ; it deflagrates and yields *metallic* silver, which may be easily fused at a stronger heat.

Experiment 4.—Add some ammonia to a solution of silver nitrate ; the dark-grey precipitate is silver hydrate, AgOH . If more ammonia is added, it is redissolved.

Silver chloride, AgCl .—*Experiment 5.*—To a solution of silver nitrate add hydrochloric acid or a solution of common salt ; you obtain a white curdy precipitate of *silver chloride*. This precipitate is so insoluble in water that it will impart a cloudiness to a solution of silver diluted a millionfold ; it is, however, easily dissolved by ammonia.

Experiment 6.—After having decanted the supernatant liquid, rub the silver chloride with a cork upon a sheet of paper, and let it dry in a dark place—in a drawer, for instance ; it remains white. Now place the sheet partly in a book, so that one-half may be exposed to the light ; this part soon acquires a violet, and finally a black colour, while that protected from the light remains white. Thus *light* is capable of decomposing this salt. On this action of the solar light on certain substances is founded the art of photography, in which light acting upon the salts of silver (especially the chloride, bromide, and iodide) may be made to yield faithful representations of any objects.

Silver sulphide, Ag_2S .—If you add sulphuretted hydrogen to a solution of silver you obtain a black precipitate of silver sulphide. This compound occurs in nature as the most important silver ore ; it is called *silver-glance*.

CHAPTER II.

METALLIC DIADS.

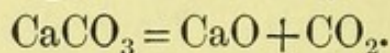
GROUP i.—Calcium, Strontium, Barium, Magnesium, Zinc.

CALCIUM.

Ca = 40.

THIS metal never occurs pure, but many of its compounds, such as chalk, marble, limestone, and gypsum, are found in enormous quantities. Calcium is a yellow metal of no practical value. It can only be separated with the greatest difficulty. It is very light (Sp. Gr. 1.6), and decomposes water at ordinary temperatures, though with less energy than potassium and sodium exhibit.

Calcium oxide (Lime), CaO.—*Experiment 1.*—Heat a small piece of chalk on charcoal before the blow-pipe for several minutes. It will be converted into calcium oxide, or *quicklime*, by losing carbonic anhydride:



Lime is obtained by burning a mixture of coal and chalk, or limestone, in chambers called *lime kilns*.

Calcium hydrate (Slaked lime), Ca(OH)₂.—*Experiment 2.*—Lime combines eagerly with water to form the hydrate. Put some lime in a saucer, and sprinkle it with water. In a few minutes, it cracks in various places, and becomes very hot; steam is copiously evolved, while the lime crumbles to a white powder, called *slaked lime*. This is calcium hydrate. It is alkaline and caustic, and slightly soluble in water. Put a little in a bottleful of water, and shake for a few minutes. A portion will dissolve. Allow the remainder to subside, and pour the clear solution into a fresh bottle, and label it *lime-water*. It becomes milky

on exposure to the air, from the formation of calcium carbonate by the absorption of carbonic anhydride, hence it is often used as a test for that gas (p. 40).

Calcium carbonate, CaCO_3 , exists in nature as chalk, marble, and limestone, and as the beautiful, transparent crystals of Iceland spar. Calcium carbonate is insoluble in pure water, but soluble, to a small extent, in water impregnated with carbonic acid. It is therefore found in most natural waters, and is a frequent cause of what is termed the *hardness* of water.

Experiment 3.—Pass a current of carbonic anhydride gas through a little lime-water in a wine-glass. The lime-water at first becomes milky, from the formation of carbonate: but after a time, the precipitate dissolves in the dissolved gas, and the liquid becomes clear. The carbonates of magnesium, iron, lead, and zinc, behave in the same way.

Experiment 4.—Boil a little of the above solution in a test-tube. The CO_2 which held the carbonate in solution is expelled, and the calcium carbonate is once more precipitated. This explains why water which owes its hardness to carbonates is softened by boiling. The white deposit which collects inside tea-kettles and boilers is formed in this way. For this reason hardness due to carbonates is called *temporary hardness*.

Experiment 5.—Temporary hardness may be removed from water without boiling. Take a little of the solution prepared in Experiment 3, and add *lime-water* to it. The lime-water will neutralise the carbonic acid which held the carbonate in solution, and the whole of the calcium will then be precipitated as carbonate. This process for softening water is called “Clark’s process.” It has been found very useful on the large scale.

Experiment 6.—Sodium carbonate may be used instead of lime-water in the preceding experiment. The sodium carbonate combines with the carbonic acid, forming hydrogen sodium carbonate, $\text{Na}_2\text{CO}_3 + \text{H}_2\text{CO}_3 = 2\text{HNaCO}_3$. We have already seen that carbonic anhydride can be expelled from calcium carbonate by strong heat, and also (p. 179) by the action of acids.

Calcium sulphate, CaSO_4 , is found native as gypsum,

alabaster, &c. It then contains two molecules of water. By heating gypsum to about 120°C . (248°F .) the water is driven off, and a white powder, the so-called *plaster of Paris*, is left. This, on being mixed to a paste with water, again unites with it, and the mass expands and becomes solid. Calcium sulphate is slightly soluble in cold water; and even less in hot, and is a frequent cause of hardness in spring water. The hard fur which accumulates in kettles and boilers is mainly composed of this salt. Hardness due to sulphates is not removed by boiling, and is therefore called *permanent* hardness.

Calcium chloride CaCl_2 .—This salt is formed when marble is dissolved in hydrochloric acid, as in the preparation of carbonic anhydride. The solution, if very much concentrated by evaporation, deposits colourless crystals ($\text{CaCl}_2 \cdot 6\text{H}_2\text{O}$); or, if evaporated to dryness, furnishes a spongy mass of the anhydrous chloride, which, in this state, is much used for drying gases, as it absorbs water with great avidity.

Calcium fluoride, CaF_2 , commonly known as fluor-spar, occurs crystallized in Derbyshire and elsewhere, and is interesting as the source of fluorine.

The remaining compounds of calcium are of little importance, with the exception of calcium chloro-hypochlorite, $\text{CaCl}(\text{ClO})$, or bleaching-powder, which has already been described under chlorine, and calcium phosphate, $\text{Ca}_3(\text{PO}_4)_2$ or bone earth, which was discussed under phosphorus. Calcium phosphate occurs native as apatite and phosphorite.

STRONTIUM.

$\text{Sr} = 87.5$.

The compounds of this metal are similar to those of calcium. Their chief sources are the carbonate called *strontianite*, and the sulphate known as *celestine* by mineralogists. The metal resembles calcium, both in appearance and the difficulty which attends its extraction. Strontium oxide, SrO , or strontia, like lime, combines with water to form a hydrate. It is most easily prepared by heating the nitrate. By dissolving the carbonate in different acids, the other salts of strontium can be prepared. The most

remarkable characteristic of the strontium salts is that of communicating a crimson colour to the flame of burning substances. Strontium nitrate, like the other nitrates, deflagrates upon burning charcoal, and is used for producing a crimson flame in fireworks. Strontium chloride is soluble in alcohol, and imparts a crimson tint to its flame.

BARIUM.

Ba = 137.

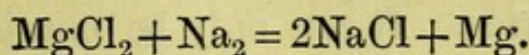
Barium compounds are more extensively distributed than those of strontium. It chiefly occurs as *heavy-spar* (sulphate) and *witherite* (carbonate). The metal can only be obtained with difficulty. Barium oxide, BaO, or baryta, is best obtained by heating the nitrate: it forms a hydrate with water which is much more soluble than the hydrates of calcium and strontium. By passing air or oxygen over the heated oxide the curious compound, barium peroxide, BaO₂, is obtained. This substance is used for preparing hydrogen peroxide. Barium carbonate is insoluble in water. Barium sulphate is insoluble in everything, and is obtained whenever a soluble barium salt is added to sulphuric acid or a sulphate. For this reason *barium chloride*, BaCl₂, and *barium nitrate*, Ba(NO₃)₂, which are soluble salts, are much employed as tests for the presence of sulphuric acid and sulphates (p. 150).

Barium salts communicate a green colour to flame: unlike the strontium compounds, they are very poisonous.

MAGNESIUM.

Mg = 24.

The chief sources of magnesium are *Magnesite* MgCO₃ and *dolomite*, or magnesian limestone, which is a carbonate of calcium and magnesium. The element occurs also in sea- and spring-water as sulphate and chloride. The metal is now prepared in somewhat large quantities, by decomposing heated magnesium chloride with sodium:



Magnesium is a silver-white metal, which does not

tarnish in dry air. Its specific gravity is 1.75. It is malleable, and may be rolled into ribbon or drawn into wire. The most notable property of magnesium is the dazzling brilliancy with which it burns when held in a flame. This fact has led to its adoption as an illuminating agent in certain cases where a strong light is required. For this purpose it is used in the form of ribbon or powder, and burnt in specially contrived lamps. The light of burning magnesium forms a valuable substitute for sun-light in photography, owing to its extreme richness in rays of chemical power. Magnesium rapidly dissolves in most acids, with the evolution of hydrogen.

Magnesium oxide (Magnesia), MgO .—*Experiment.*—Set fire to a piece of magnesium wire the flame of a spirit-lamp, or Bunsen's burner. It will burn rapidly, and be converted into a pure white powder. This is magnesium oxide, or *magnesia* (sometimes sold as *calcined magnesia*), formed by the direct union of its constituents. It is prepared in quantity by strongly heating the carbonate, which, like chalk, loses carbonic anhydride and yields the oxide. When treated with water, the oxide yields a *hydrate* analogous to that of calcium, but much less soluble in water.

Magnesium carbonate, $MgCO_3$, occurs crystallized as *magnesite*, and is the chief constituent of the *magnesia alba* of the shops. It is soluble, with effervescence, in acids.

Magnesium sulphate (Epsom salts) $MgSO_4 \cdot 7H_2O$, is the chief salt of magnesium, owing to its extensive use in medicine as a purgative. It is largely prepared by dissolving dolomite in sulphuric acid, and from its superior solubility is easily separated from the very sparingly soluble calcium sulphate which is formed simultaneously with it. To the chemist magnesium sulphate is valuable as a test for phosphates (p. 171).

Magnesium chloride, $MgCl_2$, is a very soluble salt, prepared by neutralizing hydrochloric acid with magnesium carbonate or oxide. It can only be obtained in the solid state by indirect means, as the solution decomposes on evaporation into *magnesia* and hydrochloric acid.

ZINC.

 $\text{Zn} = 65.$

Zinc resembles magnesium closely, but is more easily extracted from its ores, the chief of which are *zinc blende* (sulphide) and *calamine* (carbonate). Zinc is obtained by converting the ore into oxide by roasting in a current of air. The oxide being mixed with coal dust is then introduced into an iron retort and subjected to distillation. The carbon removes the oxygen as carbonic oxide, while the liberated zinc being converted into vapour, is carried over and condensed in a vessel below. Zinc readily dissolves in acids, with evolution of hydrogen; hence it is much employed for the preparation of that gas.

Not very long ago, zinc was hardly used except for making brass and pinchbeck; but since the art of rolling it out into sheets, of forging it, and of drawing it out into wire, has been acquired, it is used also for the manufacture of many articles which were formerly made of lead, copper, and iron; for instance, for making nails, gasometers, gas-pipes, gutters, and for roofs of houses, for lining refrigerators, as it is harder, and yet lighter than lead, cheaper than copper, and less liable than iron to be destroyed by air and water. It usually occurs in commerce in the form of plates, which are so brittle that they may be broken by the hammer into small pieces; the fresh fracture exhibits a crystalline structure, and a bluish-white colour.

Experiment 1.—If a piece of polished sheet zinc be exposed to the action of water and air, it will become gradually covered with a white film; it rusts like iron, but the rust of zinc has a white colour. In iron the oxidation proceeds rapidly towards the interior, but not in zinc, or only very slowly; therefore articles made of zinc, when exposed to the wind and weather, last much better than those made of iron, and for this reason, also, iron articles are frequently coated with zinc (galvanised iron). Zinc attracts not only oxygen, but also some carbonic acid, from the air, and this may be recognised by the effervescence which follows when some acid is dropped upon the tarnished metal.

Experiment 2.—Hold a rod of zinc by means of a pair

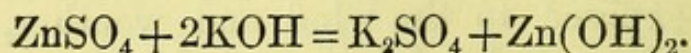
of tongs or pincers in the alcohol-flame, until it hisses if you touch it with a piece of moist wood; if you now quickly hammer it upon a stone or anvil previously heated, it does not break, but spreads out like lead into a thin, coherent sheet. Zinc has the singular property of *being ductile between 212° F. (100° C.) and 302° F. (150° C.)*, but *below or above this temperature is brittle*. Ever since it has been known that zinc is thus affected by heat, it has been found easy to overcome the difficulties which formerly opposed the conversion of this metal (which is unpliant when cold) into sheets and wire.

Experiment 3.—Zinc, when heated to 433° C. (811° F.) *melts*, as may easily be seen by holding a small piece of it in an iron spoon over an alcohol flame. In this case a grey film is formed; but this after a time assumes a yellow colour, and is converted into pure oxide (ZnO). On cooling, the yellow colour passes to white; the oxide of zinc belongs to those substances which present a colour in the heat different from the colour at the ordinary temperature.

Experiment 4.—At a still stronger heat (1040° C. = 1904° F.), zinc evaporates and burns at the same time with a blue flame, as may be seen by heating a thin strip in the blow-pipe flame. The product of the combustion is zinc oxide, ZnO, portions of which are carried up, and being very light, float in the air in flocks.

Zinc oxide, ZnO, is a white light powder prepared by heating the metal strongly in air. It is insoluble in water, but dissolves in acids, forming the zinc salts.

Zinc hydrate, Zn(OH)₂.—*Experiment 5.*—Add potassium or sodium hydrate to a soluble zinc salt, such as the sulphate. A white precipitate of zinc hydrate will be produced:



Zinc hydrate is soluble in the alkaline hydrates, therefore care must be taken not to add too great an excess, or the precipitate will disappear. By heating the hydrate, water is driven off and the oxide is obtained.

Zinc sulphate (*White vitriol*), ZnSO₄·7H₂O, is one of the most important salts of zinc. It is easily soluble and

crystallizes in colourless prisms, containing nearly half their weight of water. It may easily be obtained by evaporating the waste liquors left after generating hydrogen from zinc and sulphuric acid. In commerce the sulphate is prepared by gently roasting the native sulphide (zinc-blende) in a current of air. The sulphide takes up oxygen and is converted into sulphate.

The other salts of zinc are of little interest. The *carbonate* may be prepared by adding sodium carbonate to zinc sulphate; the *sulphide*, by adding ammonium sulphide. They are both white, insoluble substances. The native sulphide possesses a reddish colour, owing to impurities. The *chloride* is used as a disinfectant under the name of "Burnett's disinfecting fluid."

Cadmium is a somewhat scarce metal, often found in zinc ores. It is very similar to zinc, but yields a beautiful yellow sulphide when its nitrate is treated with ammonium sulphide. Artists call the sulphide *cadmium yellow*.

GROUP ii.—Copper, Mercury, Lead.

COPPER.

$\text{Cu} = 63.5.$

Copper is often met with in the free state (native copper), but seldom in any quantity. The following are some of the more important ores of the metal:—

- | | |
|--|-----------------------------|
| 1. Red oxide, Cu_2O . | } Not very abundant. |
| 2. Black oxide, CuO . | |
| 3. Copper-glance, Cu_2S . | |
| 4. Indigo Copper, CuS . | |
| Malachite, $\text{CuCO}_3, \text{Cu}(\text{HO})_2$. | } The chief Australian ore. |
| 6. Azurite, $2\text{CuCO}_3, \text{Cu}(\text{HO})_2$. | |
| 7. Copper Pyrites, CuFeS_2 . | } The chief English ores. |
| 8. Purple Copper, Cu_3FeS_3 . | |

The processes by which the metal is smelted from these ores differ very much. The ores which do not contain sulphur are sometimes reduced by fusion (after previous roasting) with coke and lime. The roasting converts the metal to oxide; this when heated with carbon loses its oxygen, while the silica, which would otherwise combine

with some of the copper, is removed by the lime. The ores which contain sulphur, and especially the English ores, require an extremely complex series of processes for their reduction.

In ancient times copper was chiefly obtained from the island of Cyprus, whence its Latin name *cuprum*. The red colour and other physical characters of the metal are well known. Its specific gravity is 8.9; rather higher than that of iron (7.8). It is very tough and *ductile*, that is, it can be drawn into fine wire. It melts at a bright red heat.

Copper is not easily altered by exposure to the air. In pure air or water it undergoes no change. In moist air it takes up CO_2 , and acquires a green film of basic carbonate. At a red heat it takes up oxygen, and yields black scales of oxide, CuO . It combines easily with chlorine, bromine and iodine, and at a red heat, with sulphur and phosphorus. It is not acted on by dilute hydrochloric or sulphuric acid. Boiled with strong sulphuric acid, it yields copper sulphate and sulphurous anhydride (p. 146). Dilute nitric acid dissolves it easily, with liberation of nitric oxide (p. 164).

Alloys of Copper.—Copper forms very important alloys with several other metals.

Gold and copper form the common gold, *silver and copper* the common silver, from which gold and silver articles and coins are made.

The well-known *brass*, and other metallic compounds having the appearance of gold, such as *tombac*, *similor*, *prince's metal*, *red brass*, &c., are composed of zinc and copper. *Spurious gold-leaf* is made by hammering out *tombac* into exceedingly thin leaves, which, when finely pulverised, constitutes the so-called *gold bronze*. *Yellow metal* contains 2 of zinc to 3 of copper. It is used for sheathing ships' bottoms. *Purple or copper bronze* is prepared by gently heating the gold-coloured bronze till it turns to a purple-red colour.

Zinc, nickel, and copper constitute the ingredients of German silver (*packfong, white copper*).

Tin and copper form a very hard grey alloy, from which statues, cannons, bells, mirrors, &c., are cast (*bronze, gun-metal, bell-metal, speculum-metal*).

Copper, like mercury, forms two distinct classes of salts, one denominated *cuprous*, the other *cupric*. Thus we have cuprous oxide, Cu_2O , and cupric oxide, CuO ; cuprous chloride, Cu_2Cl_2 , and cupric chloride, CuCl_2 .

In the cuprous compounds, two atoms of metal behave as a single diad atom (Cu_2). The cuprous compounds are, for the most part, unimportant.

Cuprous oxide, Cu_2O .—*Experiment 1.*—On heating a piece of bright copper in a smokeless flame, it passes through various shades of colour—crimson, violet, blue, and finally to grey. On speedily quenching the metal in water it becomes brownish-red; this red coating is cuprous oxide.

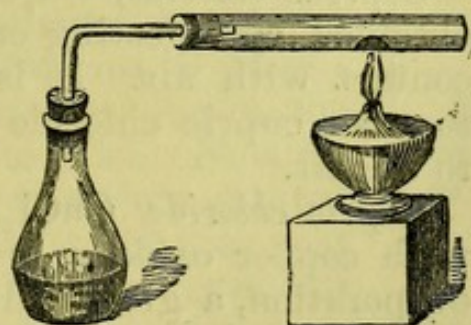
Cuprous oxide, when thrown into melting glass, colours it *blood-red*; in this manner a beautiful red colour is flashed on glass in the glass factories. This accounts, also, for the red colour of the slag which forms during the calcination and fusion of copper.

Experiment 2.—Cuprous oxide may easily be prepared by boiling a solution of copper sulphate and grape or starch sugar, to which an excess of an alkaline hydrate has been added. This will be further noticed under sugar.

Cupric oxide, CuO .—If the copper is heated for a longer time in the air it becomes coated with a black substance. This is cupric oxide; it contains a larger proportion of oxygen. By long-continued ignition, the whole mass of the copper may be converted first into cuprous and ultimately into cupric oxide. The glowing cinders which fall in the workshops of the copper-smiths (copper-scales) consist of a mixture of the two oxides.

Cupric oxide is extensively used as an agent for effecting the combustion of organic substances, with a view to their ultimate analysis. For this purpose it is usually prepared by heating cupric nitrate.

Experiment 3.—Introduce some cupric oxide into a hard glass tube, heat it, and then pass over it a stream of hydrogen, which is evolved by zinc and diluted sulphuric acid,



and then heat the tube. In the heat, the hydrogen abstracts from the oxide of copper its oxygen, and forms with it water, which escapes. A process by which this reaction has been utilized as a means of determining the composition of water has already been described (p. 135).

Cupric hydrate, $\text{Cu}(\text{OH})_2$.—*Experiment 4.*—Most of the metallic hydrates can be formed by adding an alkaline hydrate to a salt of the metal. Thus, on adding potassium hydrate to cupric sulphate there are obtained potassium sulphate and cupric hydrate. The latter substance is precipitated as a light-blue powder. Boil the liquid in which it is contained, and it will become black, for it is resolved into black cupric oxide and water: $\text{Cu}(\text{OH})_2 = \text{CuO} + \text{H}_2\text{O}$. This furnishes another example of chemical decomposition effected by mere elevation of temperature.

Experiment 5.—Repeat the former experiment, but instead of potash take ammonia; here also the copper hydrate is first precipitated, but this is redissolved by adding more ammonia, forming a splendid *blue liquid*. Ammonia is therefore a test for salts of copper. Pour upon the blue liquid an equal quantity of strong alcohol, and direct the stream against the side of the glass, so that the alcohol may float on the surface; after the lapse of twenty-four hours, a mass of dark-blue acicular crystals is perceptible, which consist of a compound of copper sulphate with ammonia, and are called *cupric ammonio-sulphate*. By dissolving them in water, the *blue liquid* of the apothecaries' show-bottles is prepared.

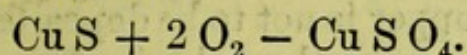
Cuprous chloride, Cu_2Cl_2 , is an unimportant compound, obtained by digesting cupric chloride with copper, out of contact with air. It is colourless, but rapidly changes to green cupric chloride when its acid solution is exposed to the air.

Cupric chloride, CuCl_2 .—By boiling hydrochloric acid with copper oxide, a green solution is obtained, and, by evaporation, a green salt—cupric chloride. It is soluble in spirit, and the solution, if ignited, burns with a beautiful green flame.

Cupric nitrate, $\text{Cu}(\text{NO}_3)_2 \cdot 3\text{H}_2\text{O}$.—Beautiful blue de-

liquescent crystals, prepared, as already described, by dissolving copper in dilute nitric acid.

Cupric sulphate (*Blue vitriol*), $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$.—This important salt is prepared in large quantities by roasting native or artificial sulphide:



It is also formed by boiling copper with sulphuric acid, sulphurous anhydride being evolved (p. 146).

Experiment 1.—If half an ounce of copper sulphate is heated to boiling with an ounce and a half of water in a porcelain basin, and then boiled a few minutes with some granulated zinc, metallic copper separates as a powder, and zinc sulphate is formed. The powder obtained is washed, and then boiled with water and a few drops of sulphuric acid, in order to remove all the zinc. It must be dried quickly, but not at a high heat, for copper in this state of minute subdivision attracts oxygen with more avidity than when it is in a compact mass.

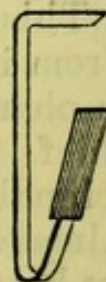
Experiment 2.—Push an iron rod into a good-sized, large-mouthed phial, forcibly enough to break out the bottom, file off the sharp edges of the fractured part, and bind some moistened bladder over the mouth of the phial. Then twist a wire firmly round the phial, in such a manner as to form two or three supports, by means of which it may be suspended in a tumbler.

Fig. 79.



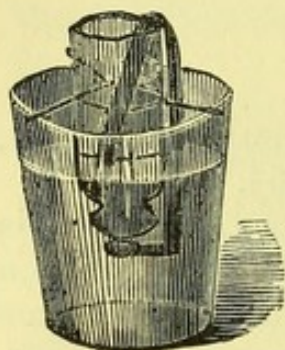
Let a strip of strong sheet zinc, of the width of the finger, and five inches long, be soldered to a strip of thin copper plate, ten inches long, and bend the strip of copper as represented in the annexed figure. Put a coin upon the lower horizontal part of a copper strip,—for instance, a bright half-crown,—or some other metallic object, the impression of which you wish to take. Now fill the phial three-quarters full with very diluted sulphuric acid (one drachm of sulphuric acid to two ounces of water), introduce the zinc, and suspend the apparatus in a tumbler, in which a saturated solution of copper sulphate, and also a few whole crystals of copper sulphate, have been put. In the course of a few

Fig. 80.



minutes the coin will be covered with a thin film of metallic copper, and after some time, with a layer several lines in thickness, which may be removed as a coherent mass.

Fig. 81.



Tallow or wax must be smeared over those parts of the coin and plate on which the copper is not to be deposited. The sunk impression thus obtained may be used in the same way again, instead of the coin, as a mould for obtaining a raised impression. When the evolution of the gas in the phial has ceased, a few drops of strong sulphuric acid may be stirred in, or the liquid, which contains zinc sulphate in solution, may be replaced by a fresh supply of diluted sulphuric acid. Salt water may also be used instead of sulphuric acid, but then the separation of the copper takes place more slowly.

The decomposition of salt has, in this case, been effected by the *voltaic current*, which is always generated when different kinds of metals, in contact, are introduced into liquids capable of being decomposed. The bladder is a porous substance, which allows the voltaic current to pass, but prevents the liquids from mixing.

This experiment is a modification of Experiment 6, page 12. It affords an illustration of the art of *electrotyping*. Solutions of gold, silver, &c., may be decomposed in the same manner (*electro-plating* and *gilding*).

MERCURY.

Hg = 200.

This metal exists at ordinary temperatures as a liquid. From its appearance it is commonly called *quicksilver*, and in pharmacy *hydragyrum*. It often occurs native; but the chief supply is obtained from the native sulphide (cinnabar). The most celebrated mine is that of Almaden in Spain. From cinnabar it is easily obtained by heating the ore in a free current of air; the sulphur burns to sulphurous anhydride and escapes, while the mercury volatilizes and can readily be condensed. In the northern regions of the earth mercury becomes solid

in winter, at a temperature of -38° F. (-39° C.); but in our climate it can only be solidified by artificial frigorific mixtures. Mercury boils at 357° C. (675° F.), and yields a colourless vapour, the specific gravity of which is 100. Hence it is easy to purify the metal by distillation.

Experiment 1.—Fasten to the cork of a phial containing mercury a piece of wood, affixing to the bottom of the latter some genuine gold-leaf; the gold, after some days, will have assumed a white colour, and be converted into an alloy of gold and mercury. It is obvious from this, that fumes of mercury must be contained in the air of the phial, and that mercury, like water, *evaporates slowly even at ordinary temperatures*. The vapour of mercury, and the preparations of mercury, are very injurious; they first produce involuntary salivation, and afterwards lingering, dangerous maladies; therefore, in experimenting with mercury, not only the inhalation of the fumes should be avoided, but it must be weighed and decanted over a bowl, so that, if any portion of it should happen to be spilt, it may not fall upon the floor. As, in comparison with water, mercury boils at a very high, and freezes at a very low temperature, and as it has a great specific gravity (13.6), it is excellently adapted to the construction of thermometers, barometers, areometers, &c.

Mercury, if quite pure, retains its metallic lustre in the air and water, and it is therefore ranked among the *noble* metals; but if it is mixed or adulterated with other metals, as lead, tin, or zinc, a grey film will gradually form upon its surface.

When mercury is long heated in contact with air, at a temperature near its boiling point, it absorbs oxygen, and is converted into the scarlet *mercuric oxide*, HgO , from which oxygen gas was first prepared by Priestley. Mercury is not acted on by hydrochloric acid, nor by *cold* sulphuric acid. When it is boiled with strong sulphuric acid, *mercuric sulphate*, HgSO_4 , is formed, and sulphurous anhydride expelled. Nitric acid, even when cold, dissolves it easily. Heated in chlorine it takes fire, and *mercuric chloride*, HgCl_2 , is formed.

Mercury, like copper, forms two series of compounds, the mercurous and mercuric. In the former two

diad atoms of the metal play the part of a single diad atom, as the following comparison of a few important compounds will show.

	Mercurous.	Mercuric.
Chloride	$(\text{Hg}_2)\text{Cl}_2$.	Hg Cl_2 .
Oxide	$(\text{Hg}_2)\text{O}$.	Hg O .
Nitrate. . . .	$(\text{Hg}_2)(\text{NO}_3)_2$.	$\text{Hg}(\text{NO}_3)_2$.

OXIDES.

Mercurous Oxide—Mercury Suboxide— Hg_2O .

Experiment 1.—Dissolve a globule of mercury in a slight excess of cold rather dilute nitric acid. Mercurous nitrate is formed, and nitric oxide set free. To the solution add caustic potash. A black precipitate of mercurous oxide is formed. It is very unstable.

Mercuric oxide.—Mercury peroxide.— HgO .—Dissolve a globule of mercury in excess of *boiling* nitric acid. Mercuric nitrate is thus obtained. On adding potash solution a yellow precipitate will be produced: it is *mercuric oxide*. Wash and dry it. This substance may also be obtained by heating solid mercuric nitrate, and by the direct union of its constituents with the aid of heat.

Mercurous nitrate, or mercury sub-nitrate, $\text{Hg}_2(\text{NO}_3)_2$.—This salt, it has already been stated, is formed by acting on mercury with cold dilute nitric acid. It may be obtained crystallized by pouring half an ounce of nitric acid and a little water on an ounce of mercury in a porcelain basin. In a few days the mercury will be covered with white crystals of the salt in question. They may be preserved and dissolved in water, with the aid of a drop of nitric acid as required.

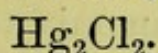
Experiment 1.—If a drop of the solution of mercurous nitrate be rubbed upon a copper coin, the mercury separates as metal, and effects a *false silvering* of the copper.

Experiment 2.—Make a stroke across a thin brass plate with a wooden stick that has been dipped in the solution of mercury; if sometime afterwards the plate is bent at this place it will break, as though it had been cut; because the reduced mercury penetrates the brass with great quickness, and renders it brittle. Thus the brazier can make use of this solution instead of shears.

Mercuric nitrate, $\text{Hg}(\text{NO}_3)_2$.—Formed in the manner already described: when heated, red fumes escape, and mercuric oxide is formed.

CHLORIDES.

Mercurous Chloride, or Mercury Subchloride (Calomel),

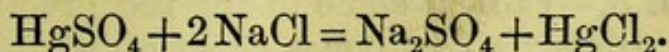


Experiment 1.—Add some hydrochloric acid, or a solution of common salt, to a diluted solution of mercurous nitrate; a heavy white precipitate of mercurous chloride is produced, which is *insoluble* in water. When well washed and dried, this salt of mercury forms the highly important medicine known as *calomel* (*precipitated*). If some of it is moistened with an alkaline hydrate it becomes black, from the formation of mercurous oxide; thus is explained the Greek name *calomel* (καλός, beautiful, μέλας, black). Calomel is now almost exclusively obtained by a process which will be described presently.

Mercuric Chloride (Corrosive Sublimate), HgCl_2 .

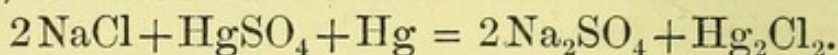
Experiment 1.—Heat some mercuric oxide with hydrochloric acid, and continue adding the latter till a complete solution is obtained; the white prismatic crystals which separate on cooling are mercuric chloride, or mercury bi- or perchloride—one of the *most violent poisons*. The same compound is obtained on a large scale, in white, transparent, heavy masses, by the following process:

Experiment 2.—Grind sixty grains of mercuric sulphate in a mortar, with an equal quantity of common salt. Heat half the mixture thus obtained in a test-tube. A white crystalline crust will condense on the upper and cooler part of the tube, and may be removed and dissolved in boiling water. It consists of mercuric chloride:



Experiment 3.—Add twenty grains of mercury to the other half of the mixture, and grind them together until the mercury loses its fluidity and becomes incorporated with the other materials to a grey powder. Heat this mixture in a tube as before. A similar white crust is

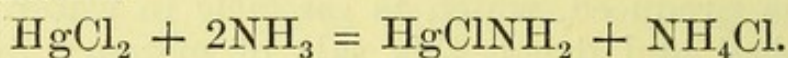
produced; but it cannot be dissolved. It is *mercurous chloride*, or calomel:



The same result would be produced by heating the proper proportion of mercury directly with mercuric chloride instead of the materials for preparing it.

Experiment 4.—Add ammonia to a solution of mercuric chloride; a white and not a yellow precipitate is formed.

This is called white precipitate or mercuric amido-chloride. It contains the monad radical NH_2 , derived from ammonia.



Mercuric iodide, HgI_2 .—This deadly poison is chiefly interesting from its brilliant red colour, and the change which the latter undergoes under the influence of heat. It is very volatile, and its vapour is remarkable as being more than fifteen times heavier than air (Sp. Gr. 227, $\text{H} = 1$). It may be prepared by the direct union of its constituents.

Experiment 1.—Triturate a few grains of iodine in a mortar with a small globule of mercury and a drop of alcohol. The formation of the iodide will become apparent by its red colour.

Experiment 2.—By far the best method of forming this body is by the addition of an iodide to a soluble mercuric salt. Add potassium iodide slowly to mercuric chloride. Each drop of the iodide as it falls and mixes with the other solution presents the most beautiful appearance. Festoons of a canary-yellow precipitate are formed, and soon change their colour to a brilliant red. The colour disappears on stirring if the mercuric chloride be in excess, but is reproduced by the addition of more iodide. The red precipitate is also soluble in excess of potassium iodide.

Experiment 3.—Collect some of the red precipitate, wash and dry it. Smear a little on a sheet of white paper, and warm the latter over a lamp. The red colour will be changed to yellow. Rub the changed colour with a hard body as a stirring-rod; wherever the rod touches the red colour will reappear. The original colour also returns spontaneously in a few days. The difference in colour is

not owing to altered composition, but to a change in molecular arrangement.

Mercuric sulphate, HgSO_4 .—This salt is obtained by boiling mercury with sulphuric acid in a Florence flask until a white mass is left. It is chiefly used for preparing calomel and corrosive sublimate. It is decomposed by water.

Mercuric Sulphide, HgS .

Experiment.—If a solution of mercuric chloride be agitated with a little sulphuretted hydrogen water, or ammonium sulphide, a white precipitate is formed, which, on adding more of the precipitated body, becomes yellow, brown, and finally black; the black substance is mercuric sulphide. This compound is also obtained by mixing mercury with melted sulphur, or indeed by rubbing it for a day with flowers of sulphur (*Ethiops mineral*). If this black sulphide is sublimed in a glass tube, then a blackish-red crystalline mass is obtained, the colour of which, by friction, passes into the most magnificent scarlet red. The sulphide in this state is called *vermilion*, or cinnabar. The red and the black sulphide have precisely the same composition, and yet a very great difference in appearance. Vermilion is also frequently prepared in factories in the moist way, by triturating together for a day mercury, sulphur, and a solution of potassa. When vermilion is pure, it volatilises completely on a glowing coal, emitting, at the same time, a blue sulphurous flame; but if adulterated with red lead, beads of metallic lead remain behind. On account of its insolubility, it is less prejudicial to health than many of the other compounds of mercury.

Cinnabar also occurs in nature, and we have it in the most important ore, from which we obtain mercury on a large scale. Small globules of pure mercury are also found in many porous stones.

AMALGAMS.

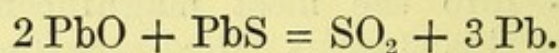
Experiment.—Introduce a globule of mercury into a porcelain dish, put upon it a piece of lead, and let them remain for some time in contact; both metals will in-

timately combine together. If the proportion of mercury is small, a friable mass is produced; but by increasing the quantity, a paste, and, if still more is added, a liquid solution, is obtained. Mercury will combine in a similar manner with many other metals, forming what are called *amalgams*. The *amalgam of tin* is especially important for silvering glass.

LEAD.

Pb = 207.

Lead does not occur free in nature, but chiefly as sulphide, PbS, called by mineralogists galena. In smelting galena, the ore is first roasted in a current of air. During the process part of the sulphide combines with oxygen, and there are formed lead oxide and sulphurous anhydride, which last escapes as gas. The air being then excluded and the heat raised, the sulphide which remains reacting with the oxide, forms sulphurous anhydride and metallic lead:



Galena generally contains silver in varying quantities. The silver is economically extracted by the process already described (page 213).

The physical properties of lead, its blue lustre, its easy fusibility, softness and pliability, its high specific gravity (11.4), &c., are well known. It contracts during solidification, and it is therefore impossible to obtain sharp casts with it.

Lead-shot.—Lead may be granulated, like zinc, by being melted and poured into water. In the manufacture of *shot*, the drops of lead are let fall from such a height that they solidify before reaching the water. For making the largest-sized shot, a tower of at least 150 feet high is required. A small quantity of arsenic is usually added to the lead, to render the drops perfectly globular. As lead and arsenic are both poisonous, shot should be used with caution in washing out bottles.

Lead is not changed by exposure to air which is *perfectly free from water*, or water *perfectly free from dissolved air*. But in ordinary air it tarnishes rapidly, and in ordinary

water its surface is soon converted into *lead hydrate*, $\text{Pb}(\text{OH})_2$, which dissolves in the water, and communicates to it poisonous properties. This corrosion is, however, almost immediately arrested in many cases by the action of the carbonates and sulphates so often found in drinking water. Carbonates convert the soluble lead hydrate into an insoluble double salt ($\text{Pb}(\text{OH})_2, \text{PbCO}_3$), and sulphates into an equally insoluble lead sulphate (PbSO_4), either of which soon covers the metal with a coating which prevents the further action of the water. New lead pipes and cisterns are, therefore, more dangerous than old ones. Neither should in any case be used for water which does not contain notable quantities of carbonates, sulphates (or phosphates, which have a similar effect). Chlorides and nitrates have a contrary effect, and increase the corrosive action of the water.

Experiment 1.—To detect the presence of minute traces of lead in water, evaporate half a pint of the water to dryness, add a little nitric acid, evaporate again, very gently, till the residue is pasty, add half an ounce of distilled water, warm and filter. Through the cold filtered liquid pass sulphuretted hydrogen gas for some minutes. A very minute trace of lead would in this case yield a brown coloration; a larger quantity of the metal would cause a black precipitate (PbS).

Experiment 2.—The action of water on lead may be tried by placing bright slips of lead in two glasses, one filled with rain, the other with hard spring water, and allowing them to remain for some days. The water in each glass may then be *carefully* filtered and tested in the manner above described. It will generally be found that the spring water does not hold any lead, or only a very small trace, in solution, while the rain water has dissolved a good deal.

COMPOUNDS.

Lead oxide, PbO .—*Experiment 1.*—If lead is heated before the blow-pipe in the outer flame, it melts at 617°F . (325°C .), and is thereby coated with a grey film; indeed, it is finally entirely converted into a grey powder. This may be regarded as a mixture of oxide of lead with metallic

lead. By continued blowing, this grey colour is changed to yellow; the yellow body is lead oxide (PbO). At a stronger heat the oxide melts, and solidifies on cooling into a reddish-yellow mass, composed of brilliant scales, the well-known *litharge*. By directing upon it the *inner* blow-pipe, metallic lead will again be obtained. This easy reducibleness, which is peculiar to almost all salts of lead, together with the *incrustation of yellow oxide, deposited upon the charcoal*, is a certain test for the presence of lead.

Litharge has many applications in the arts and trades. Lead-glass (flint-glass), lead-glaze, and sugar of lead are prepared from it; the manufacturing chemist likewise prepares from it red lead, white lead, and other lead colours, and lead salts; the apothecary compounds lead plaster, by boiling it with olive-oil; the cabinet-maker makes a varnish that dries rapidly, by boiling it with linseed oil, &c.

Lead di-oxide, PbO_2 .—*Experiment 2*.—If you heat some red lead gently in nitric acid for a few minutes, it is converted partly into nitrate, which dissolves, and partly into di-oxide, which remains undissolved as a dark-brown powder.

Red oxide of lead.—*Experiment 3*.—Heat in a ladle one drachm of litharge and a quarter of a drachm of potassium chlorate; the yellowish mixture smoulders to a red powder, which must be well washed with water. The same thing happens on heating the litharge for a day, a little below melting point; and at the same time frequently stirring it. By both methods the litharge receives one-third more of oxygen; in the former case from the chlorate, in the second case from the air; and is thereby converted into Pb_3O_4 ; this compound is called *red lead*, or *minium*, and is much used as a scarlet pigment.

It may be regarded as a compound of 2PbO , and PbO_2 .

Lead hydrate, $\text{Pb}(\text{OH})_2$.—This substance is obtained as a white precipitate when potassium or sodium hydrate is added to a soluble lead salt, as the nitrate. It is soluble in excess of the alkali, and in most acids.

Lead nitrate, $\text{Pb}(\text{NO}_3)_2$.—The best solvent of lead is warm nitric acid, diluted with water; the product is lead nitrate. It may be obtained in crystals, by evaporating and cooling

the solution. Litharge may be used instead of metallic lead.

When dry lead nitrate is heated in a test tube, nitric peroxide, NO_2 , escapes and PbO remains.

Lead chloride, PbCl_2 .—*Experiment 4*.—Heat to boiling a little litharge, with hydrochloric acid and water, and decant the clear liquid from the sediment into a glass vessel; you obtain, on cooling, lustrous white acicular crystals of lead chloride (horn-lead). This salt is but very sparingly soluble in water.

Experiment 5.—If two grains of litharge and fifteen grains of sal ammoniac are fused together in an iron spoon, there is obtained a combination of a small quantity of lead chloride, with a large proportion of lead oxide, in the form of a brilliant, yellow, laminated mass, which when triturated yields a beautiful yellow powder. This powder is used by painters under the name of *Cassel* or *mineral yellow*.

Lead acetate, $\text{Pb}(\text{C}_2\text{H}_3\text{O}_2)_2 \cdot 3\text{H}_2\text{O}$, which contains one-seventh of its weight of water of crystallization, forms, the most important soluble salt of lead, *sugar of lead*, which commonly crystallizes in four-sided prisms.

Fig. 82.



Basic lead acetate is prepared by digesting a solution of sugar of lead with oxide of lead, whereby part of the oxide of lead is dissolved. This combination is kept in the apothecaries' shops in a liquid form, under the name of *solution of subacetate of lead*, or *Goulard's extract*. When mixed with water it forms the so-called *Goulard water*, which has often a milky appearance, because some lead carbonate is formed and separated by the carbonic acid of the water or air.

Lead sulphate, PbSO_4 .—This salt is easily formed when dilute sulphuric acid or a soluble sulphate is added to a solution of lead. Even in a solution of lead more than a thousand times diluted, a white turbidity is produced, since the lead sulphate is a very insoluble salt; we have, accordingly, in sulphuric acid, a very delicate test for salts of lead.

Lead carbonate, PbCO_3 .—*Experiment 6*.—Add to a solution of sugar of lead a solution of sodium carbonate as

long as a precipitate is formed; the precipitate is lead carbonate. The pigment known under the name of *white lead* is likewise carbonate of lead, but mixed with variable quantities of lead hydrate. This is prepared on a large scale in different ways. By the oldest, the *Dutch method*, a large number of jars, in which some acetic acid has been poured, are arranged in a building upon a layer of spent tan, and rolls of sheet-lead are then suspended in the jars above the acid, and the whole covered with another layer of tan. After the lapse of several months the rolls of lead are found to be mostly, if not entirely, converted into white lead.

Lead carbonate is extensively used as a white paint, and for this purpose it is very valuable, but from its poisonous nature is often productive of serious disease to those who work with it. Many less dangerous substitutes have consequently been proposed, but none have been found to equal it in opacity, or "body."

Powdered barium sulphate (heavy spar) is commonly used to adulterate inferior kinds. This impurity may be detected by dissolving the white lead in dilute nitric acid, when the barium sulphate, being quite insoluble, remains behind.

On heating white lead, carbonate anhydride is evolved and lead oxide remains.

Lead Tree.—*Experiment 7.*—Dissolve half an ounce of acetate of lead in six ounces of water, filter the liquid, pour it into a phial, and then suspend in the latter a zinc rod, by attaching it to a cork; the zinc is soon covered with a grey coating, from which brilliant metallic spangles will gradually shoot forth, finally filling up the interior of the phial. They consist of pure lead (*the lead-tree*). After twenty-four hours, no trace of lead can be found in the solution; it has



been replaced by zinc; the stronger zinc has forced the weaker lead from its combination.

This experiment might be made to illustrate not only the relative combining *powers* of the two metals, but also their equivalents. It would only be necessary to weigh the lead formed, and the zinc before and after the experiment, to ascertain that the weight of the precipitated lead

is to the loss of zinc as 207 is to 65. An atom of lead is therefore replaced by an atom of zinc. A similar experiment, in which silver is precipitated by means of copper, forming the silver-tree, was described in page 214.

Lead tartrate, $\text{PbC}_4\text{H}_4\text{O}_6$.—*Experiment 8.*—Make a strong solution of lead acetate, and add to it a solution of tartaric acid. The white precipitate of lead tartrate formed is collected on a filter, washed well with water and dried. It is interesting from the fact, that on being heated strongly in a close vessel it is decomposed, and furnishes a mixture of finely divided metallic lead and carbon which, on exposure to air, spontaneously ignites. Such a substance is termed a pyrophorus. Lead pyrophorus is prepared as follows: Select a glass tube, as thick as a lead pencil, and seal one end with the blow-pipe. Introduce sufficient dry lead tartrate to fill it for about an inch and a half; then at three inches from the closed end soften a part of the tube in the blow-pipe flame, and draw it out so as to constrict the tube to a narrow channel. When the tube is cold, hold it horizontally and shake the tartrate, so that it occupies the whole space to the constriction, and leaves a clear channel above the substance. The tube must then be heated, gradually and progressively, from the sealed end to the constriction, with a lamp. As soon as the tartrate is heated it begins to decompose and blacken, coloured drops of moisture are formed and driven along the tube, and a thick smoke is evolved which smells like burning sugar. When the evolution of fumes ceases, the heating is discontinued and the tube allowed to cool; the black substance obtained is the pyrophorus, and when shaken out into the air it takes fire, and burns with a great glow and much yellow smoke. By applying the blow-pipe flame, the tube may be sealed at the constriction, and the contents will then retain their pyrophoric powers for an indefinite time. The yellow powder produced by the ignition is lead oxide.

Lead sulphide, PbS .—This is a black powder, obtained whenever a lead salt is precipitated with sulphuretted hydrogen. It occurs native as galena, and is easily recognised by its greyish-black colour, its shining metallic lustre, and its high specific gravity.

CHAPTER III.

METALLIC DIADS, OR TETRADES.

GROUP i.—Iron, Chromium, Manganese, Aluminium,
Cobalt, Nickel.

THE first three metals of this group yield two well-defined series of compounds, analogous to those of copper and mercury. In one series the metal is a simple diad. Such compounds are sometimes called *proto-compounds*, but are more frequently designated by the termination *ous*. Ferrous chloride, or iron protochloride, $\text{Fe}^{\text{II}}\text{Cl}_2$, is an example. In the other series two atoms of the metal are generally supposed to play together the part of a single *hexad* atom, *e.g.* $(\text{Fe}_2)^{\text{VI}}$. This is perfectly consistent with the tetravalency of the metal, as will be fully explained in the introduction to organic chemistry (Part III., Chap. I.). Compounds of this kind have the termination *ic*, but are also known as *per*, or *sesqui* compounds. Ferric chloride (iron perchloride, or iron sesquichloride), $(\text{Fe}_2)^{\text{VI}}\text{Cl}_6$, illustrates their composition. They may, however, be regarded as compounds in which the metal is a triad. Ferric chloride, then becomes $\text{Fe}^{\text{III}}\text{Cl}_3$, and ferric hydrate $\text{Fe}^{\text{III}}(\text{OH})_3$. The following table of a few important iron compounds will explain the constitution of the two series more clearly than words:

Ferrous.	Ferric.	
FeCl_2	Fe_2Cl_6	Chloride.
FeO	Fe_2O_3	Oxide.
$\text{Fe}(\text{OH})_2$	$\text{Fe}_2(\text{OH})_6$	Hydrate.
$\text{Fe}(\text{NO}_3)_2$	$\text{Fe}_2(\text{NO}_3)_6$	Nitrate.
FeSO_4	$\text{Fe}_2(\text{SO}_4)_3$	Sulphate.

Besides these compounds, in which the metal is basic, several of the metals of this group (notably chromium

and manganese) yield interesting and important acid radicals. With the exception of aluminium, neutral oxides of all of them are known. No *aluminious* and no *nickelic* compounds (except nickelic oxide, Ni_2O_3) are known.

IRON.

$\text{Fe} = 56.$

This most important metal is rarely found native. The *meteorites* which sometimes fall from the sky often contain more than 90 per cent. of iron, together with nickel, cobalt, and other elements. The following are the most important ores of iron:

Fe_3O_4 , *Magnetite*, *Loadstone*.—Found in Norway, Sweden, the United States, &c. The most valuable ore of iron.

Fe_2O_3 , *Red hæmatite*.—Found in England and elsewhere. The same oxide occurs in combination with water, as *brown hæmatite*, $2\text{Fe}_2\text{O}_3 \cdot 3\text{H}_2\text{O}$.

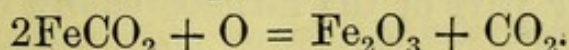
FeCO_3 (Ferrous carbonate).—Nearly pure as *spathic ore*, found in Styria, &c. *Clay-iron ore*, the great English ore, is ferrous carbonate, with clay and other impurities. It occurs in Staffordshire, &c., in grey nodules. *Black band*, the chief Scotch ore, is clay-iron, with bituminous matter.

FeS_2 , *Iron pyrites*.—Very abundant. It yields bad iron, but is valuable as a source of sulphur. It is often used as a substitute for sulphur in making sulphuric acid.

METALLURGY OF IRON.—CAST IRON—SMELTING.

The methods adopted for the separation of iron from its ores vary a good deal, according to the nature of the ores. The most important is that used in England for clay iron.

1. The ore is *roasted* at a dull red heat. This is effected by piling it in heaps, with alternate layers of coal, and burning the coal with free access of air. The ferrous carbonate loses CO_2 , just as chalk does in a lime kiln, and at the same time takes up oxygen:

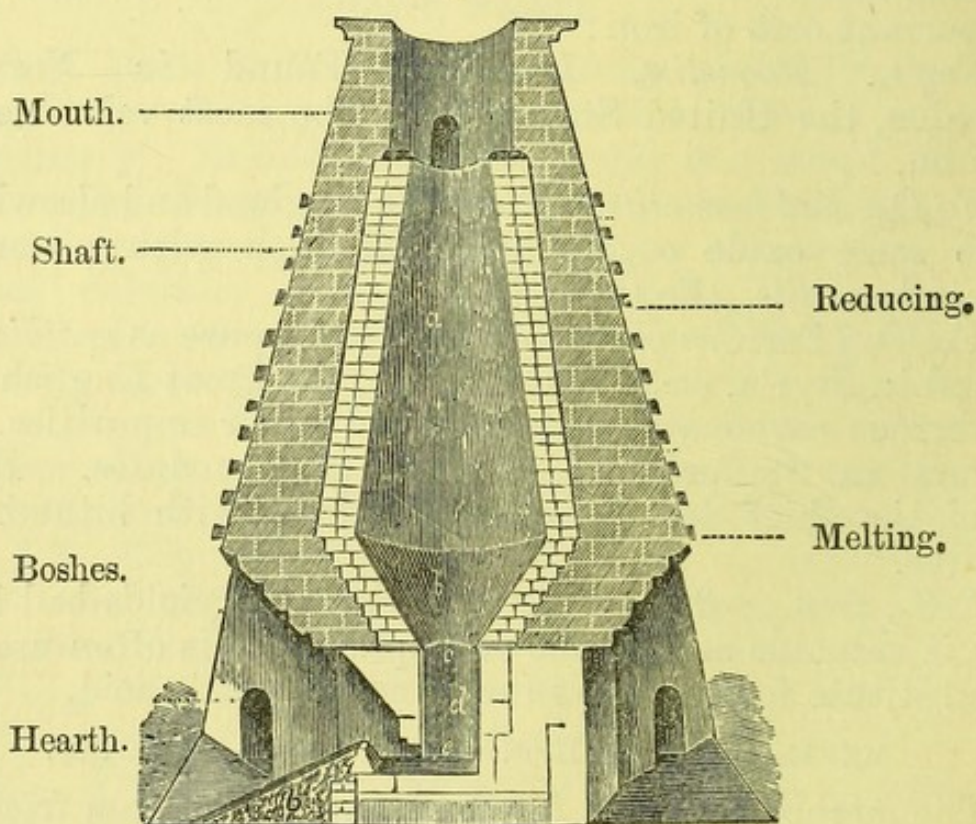


The water of the ores is expelled at the same time.

2. The *smelting* is effected in a *blast furnace* (Fig. 84),

which is from 50 to 90 feet in height. The furnace is charged with alternate layers of coal, roasted ore, and limestone (CaCO_3), and fresh charges of each are added every few hours, as the mass sinks in the furnace. In this way the action of the furnace is rendered continuous, and it can be worked for years without cessation. As the ore sinks down, the ferric oxide is reduced to metallic iron, which, in the lower and hotter part of the furnace, is melted. Melted iron accumulates on the *hearth* of the furnace, and is drawn off from time to time into rough

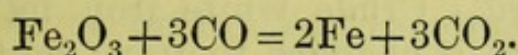
Fig. 84.



moulds printed in sand. When the iron is cold it is broken up into bars, which are called *pigs*. The combustion of the fuel is maintained by air forced in through conical apertures, called *tuyeres*, just above the hearth. To avoid waste of heat this air is now generally heated to from 350° to 700° C. (662° to 1292° F.), by the waste gases from the upper part of the furnace before it is blown in. (Hot blast.)

Chemistry of the blast-furnace.—The accompanying wood-cut represents one form of the blast-furnace, and the

technical names of the different parts of it. In the lower part of the furnace, where the air comes in contact with the fuel, the heat is most intense. Combining with the oxygen of the air, the carbon of the fuel is here converted to carbonic anhydride, which passes upward through the hot mass above. But we have already seen (p. 182) that when carbonic anhydride passes over red-hot carbon it is reduced to carbonic oxide: $\text{CO}_2 + \text{C} = 2\text{CO}$. The latter gas is, therefore, formed in large quantity. In the upper and cooler part of the furnace this carbonic oxide comes in contact with the heated ferric oxide, and at once reduces it to a porous mass of metallic iron:



As this porous iron descends into the intensely hot portion of the furnace (the boshes), in which the temperature rises to 1400°C . (2552°F .), it combines with carbon, and melts, and the liquid compound accumulates on the hearth, whence from time to time it is drawn off into the moulds. (Cast iron.)

The use of the lime in this operation is very interesting. It combines with the silica and the other foreign matters of the ore and fuel, forming with them a kind of rough glass (the *slag*), which melts at a lower temperature than iron. The melted metal is always covered with a layer of this liquid slag, which is constantly flowing off through apertures arranged for the purpose.

The slag from the blast-furnaces has generally a green or blue colour, owing to the oxides of iron and of manganese dissolved in it. It is sometimes formed into square blocks, and used for building stones.

Cast or Crude Iron.—The metal obtained by the above process is by no means pure iron, but contains 5 or 6 per cent. of carbon and other impurities, of which silicon, sulphur and phosphorus are the most important. Cast-iron, thus obtained, is characterised by the following properties:

a.) It is *fusible* at a much lower temperature than pure iron; therefore it is especially adapted for those iron articles which are made by casting. For remelting iron on a small scale, crucibles are made use of; but on a

large scale shaft-furnaces, or the so-called cupola-furnaces.

b.) Cast-iron is *brittle*, and can neither be forged nor welded (bar-iron and steel may be bent, forged, and welded). The application of cast-iron must, therefore, be limited to the manufacture of such articles as are not exposed to being bent, or to strong concussions.

There are two kinds of cast-iron in commerce, known as grey and white iron. The *grey iron* is almost black, has a granular texture, and admits of being filed, bored, &c.; the *white iron*, on the contrary, is of a silvery whiteness, possesses a lamellar-crystalline texture, and is exceedingly hard. Crude white iron, by remelting and very slow cooling, is changed to grey; on the other hand, the grey is changed to white iron by being heated and suddenly cooled. Grey iron is best adapted for castings; white iron is the most suitable for the manufacture of bar iron and steel. When grey iron is dissolved in dilute acids a mass of graphitic carbon remains, whereas in white iron almost all the carbon is in a state of combination. *Mottled iron* is intermediate between white and grey.

Malleable, Wrought, or Bar Iron.—Cast-iron, by being deprived of its carbon, is converted into *malleable iron*, and acquires the following very important properties:

a.) Bar-iron possesses great *ductility* and *tenacity*, and may be hammered or rolled into sheets, and drawn out into fine wire, which is not the case with cast-iron.

b.) At a less degree of heat than that of fusion it becomes *soft*, like wax or glass, so that two glowing pieces may be welded into one. Upon this property depends its capacity of *being welded*.

c.) Wrought-iron is sufficiently soft to be worked by steel instruments, and it does not become *harder* if, when heated to redness, it is suddenly quenched in water (steel is thereby rendered brittle).

d.) Wrought-iron is distinguished, moreover, from cast-iron by its *fibrous texture*, composed, as it were, of threads incorporated together; while cast-iron has the appearance of being a baked *granular* mass. But it is a very striking fact that fibrous wrought-iron, by repeated jolts or blows, becomes gradually granular and brittle, as, for example,

in the axletrees of carriages. By thoroughly heating and reworking, the former strength and flexibility, as well as the fibrous texture, is restored to the iron.

Wrought-iron is not entirely free from carbon; it contains, however, only 0.1 to 0.5 per cent of it. Iron entirely free from carbon is softer and more tenacious than bar-iron. Thus we see that it is the chemical combination of the carbon with the iron, as in cast-iron, which destroys these two properties of softness and tenacity.

Refinery of Iron.—The method which is employed for separating carbon from the cast-iron is very simple. The carbon is burnt out by heating the iron to fusion, and constantly stirring it (*puddling*) while exposed to a current of air (or, as in the famous Bessemer process, by forcing air through the melted metal), the oxygen of which combines with the carbon, forming carbonic oxide gas. During the operation, a considerable portion of the iron is converted by oxidation into oxide, which fuses with the sand, that is purposely strewed upon the hearth, and forms with it a heavy black slag of iron silicate. The iron mass becomes gradually more tenacious, since the iron melts with more difficulty the less carbon it contains; and finally, in the form of a loosely coherent mass (the *bloom*), is placed under a steam hammer, by a few blows of which the remaining slag is pressed out, and the iron particles are formed into a compact mass. The latter is afterwards usually hammered or rolled into bars or bands. This method of converting brittle cast-iron into ductile and malleable iron is called *puddling*. It is sometimes preceded by a separate process called *refining*.

Steel.—*Steel* holds a middle place between cast and wrought iron, both as to the quantity of carbon it contains, and other properties.

a.) If quenched when heated to redness, it is rendered hard and *brittle* (like cast-iron); if cooled somewhat more slowly, it is rendered *elastic*; and if cooled very slowly, it is soft, *ductile*, and *malleable* (like bar-iron).

b.) It is less fusible than cast-iron, and more so than bar-iron.

c.) It contains combined carbon generally less than 1 per cent.

To these properties steel owes its importance as a material for thousands of articles, especially for cutting instruments, since it may be made soft or hard, elastic or brittle, at pleasure. The article manufactured is usually first heated to redness, then suddenly cooled by quenching it in water, and afterwards *tempered*, in order to diminish its hardness and brittleness.

Experiment.—Hold a steel knitting-needle in the flame of a spirit-lamp till it is red-hot, and then quickly plunge it in cold water; it thereby becomes so brittle as to break on any attempt to bend it. Again hold the needle in the fire, and observe the changes of colour which it passes through; it will first become yellow, then orange, crimson, violet, blue, and finally dark-grey. The change of colour is owing to the formation of a film of oxide, which at first is *thin*, and has a *yellow* appearance, but gradually it becomes thicker and also darker, as the heat increases. A definite degree of hardness and elasticity of the steel corresponds to each of these tints, the needle when covered with the yellow film being the hardest and most brittle, and when presenting a blue aspect being in its softest and most elastic condition. The workmen in steel impart to their articles various degrees of hardness and elasticity by tempering; files and razors are made very hard and brittle,—saws, watchsprings, &c., soft and elastic.

Steel may be prepared in various ways:—

1. By the process of cementation, which consists in filling an iron box with bar-iron and powdered charcoal, and then maintaining the whole for several days at a red heat. The carbon gradually penetrates into the iron, thus converting it into steel (*blistered steel*).

2. In the Bessemer process, by blowing air through melted cast-iron until the impurities are oxidised and then adding a definite quantity of melted white cast-iron (*spiegeleisen*). The cast-iron supplies exactly the quantity of carbon necessary for the formation of steel.

Iron, nickel, and cobalt, are the only metals which are attracted by the *magnet* in a sensible degree. Magnetism

immediately vanishes from bar-iron when it is removed from the magnet; while steel, on the contrary, retains the magnetic power, but loses it on being heated to redness.

Absolutely pure iron can only be prepared with difficulty. Its specific gravity is 7.8. When treated with acids, pure hydrogen is evolved, whereas ordinary iron yields impure hydrogen, as is sufficiently proved by its disagreeable smell.

COMPOUNDS OF IRON.

Ferrous oxide.—*Protoxide of iron*, FeO , is prepared with difficulty, so readily does it take up oxygen and pass to a higher oxide.

Black, or magnetic oxide Fe_3O_4 .—*Experiment 1.*—Place a few grains of iron filing upon a piece of charcoal, and heat it for some minutes in the outer flame of the blow-pipe, directed upon one spot; it becomes red-hot, and the heat spreads throughout the whole mass, as is apparent from the iridescent tints which precede the red heat. The iron on cooling acquires a darker, almost a black colour, and bakes into a coherent mass of this intermediate oxide. It is the same compound as is formed when iron burns in oxygen or air (p. 127).

Fig. 85.

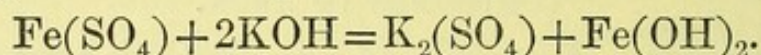


Ferric oxide.—*Peroxide of iron*, Fe_2O_3 .—*Experiment 2.*—Iron cinders (Fe_3O_4) when exposed for a long time to the outer or oxidising blow-pipe flame, become covered with a red pulverulent coating; they take yet more oxygen from the air, and become *peroxide of iron*.

It may be prepared more easily in the following manner: Place a crystal of green vitriol upon charcoal, and heat it until it has become of a brownish red colour. The salt is decomposed and the peroxide remains. The red colour of the latter is more clearly brought out by rubbing it on paper with the nail. In the same manner,

peroxide of iron remains behind when green vitriol is heated in the preparation of Nordhausen sulphuric acid; this forms an article of commerce under the name of *colcothar*, or *rouge*, and is a favourite and cheap pigment for varnish, and is also used in the polishing of glass and metals.

Ferrous hydrate, $\text{Fe}(\text{OH})_2$.—*Experiment 3*.—Add potassium hydrate to a freshly prepared solution of green vitriol (ferrous sulphate). A green precipitate of impure ferrous hydrate will be produced by double decomposition:

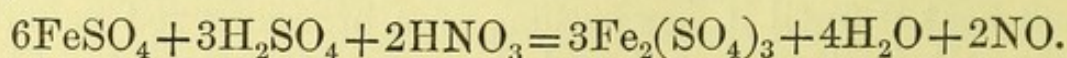


If it were pure ferrous hydrate the precipitate would be white, but is never obtained so except when oxygen is completely excluded. The green precipitate, on exposure to air, becomes brown by conversion into ferric hydrate.

Ferric hydrate, $\text{Fe}_2(\text{OH})_6$.—Ferrous salts, when kept for any length of time in contact with air and water, take up oxygen, and become ferric salts. The former are green, the latter brownish-red.

Experiment 4.—Make a solution of ferrous sulphate, filter, and set it aside. The solution will gradually become turbid, and a brown precipitate will settle on the sides of the vessel, while in a few days the bright green colour of the solution will give place to a decided brown. The change of colour indicates the conversion of the ferrous into the ferric salts.

Experiment 5.—To a fresh and dilute solution of ferrous sulphate in a test-tube add a few drops of sulphuric and nitric acids, and heat. The mixture will become of a dark brown, almost black colour, which, as the heat approaches boiling-point, suddenly vanishes and leaves the reddish colour of the per-salt. The conversion of the ferrous into ferric sulphate, which in the former experiment requires an inconvenient length of time, is in this way completed almost instantaneously by the oxygen supplied by the nitric acid:



The dark substance first formed is a compound of ferrous sulphate with nitric oxide, $2\text{FeSO}_4 \cdot \text{NO}$.

Experiment 6.—Add ammonia or potash to the solution of ferric sulphate obtained in the last experiment. Instead of a green, a red precipitate will be obtained, indicating the change which has taken place. The red precipitate is ferric hydrate, $\text{Fe}_2(\text{OH})_6$. The precipitate may be washed on a filter. Any ferric salt may readily be prepared by dissolving the moist hydrate in the acid. When the hydrate is heated ferric oxide is obtained.

Ferrous chloride, FeCl_2 , is a green salt obtained in crystals by dissolving iron in hydrochloric acid and evaporating the solution.

Ferric chloride, Fe_2Cl_6 .—This is one of the most important of the per-salts of iron. It may be obtained by dissolving ferric hydrate or ferric oxide in hydrochloric acid, or by converting the proto-chloride by boiling with nitric and hydrochloric acids.

The salt cannot be obtained crystallised, but only as a brown mass by evaporation to dryness.

Ferrous sulphate.—*Green vitriol.*—*Copperas*, FeSO_4 .—This most important salt may be obtained in a variety of ways. (1) By dissolving iron in sulphuric acid. (2) It remains in solution in the bottle in which sulphuretted hydrogen is generated by the action of sulphuric acid on ferrous sulphide (p. 143). It is manufactured on the large scale by the gradual oxidation of iron pyrites by the action of moist air, and subsequent solution and crystallization. This salt is largely used in dyeing for the production of black colours, and also in the manufacture of writing-ink. The crystals of ferrous sulphate have the formula $\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$.

Ferric sulphate, $\text{Fe}_2(\text{SO}_4)_3$.—The preparation of this salt, by boiling ferrous sulphate with nitric acid, has already been described. It may also be formed by dissolving ferric oxide, or hydrate, in sulphuric acid.

Ferric nitrate, $\text{Fe}_2(\text{NO}_3)_6$, is obtained by adding iron filings to warm dilute nitric acid as long as they continue to dissolve in it. This solution is of a brown colour, and is used in dyeing. If some nitric acid is dropped upon cast-iron, steel, or bar-iron, black spots are produced, because the iron, but not the carbon, is dissolved. These spots are darker in cast-iron and lighter in bar-iron.

Ferric acetate, $\text{Fe}_2(\text{C}_2\text{H}_3\text{O}_2)_6$, may be prepared directly, by dissolving freshly precipitated and still moist ferric hydrate in acetic acid. When the shoemaker pours beer upon iron nails to prepare the iron-black with which he blackens his leather, he obtains this salt, for, on exposure to the air, the beer is changed into vinegar.

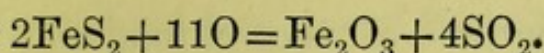
Ferrous sulphide, FeS .—On adding some sulphuretted hydrogen water to a slightly acidified solution of ferrous sulphate, no precipitate is produced; but if ammonium sulphide is added to the neutral solution, a deep black precipitate is formed; this precipitate is ferrous sulphide.

Ferrous sulphide is largely used by the chemist for the preparation of sulphuretted hydrogen (p. 143). It is best prepared in the following manner:

Experiment 7.—Heat a clay crucible to redness in a clear fire and throw in, in small portions at a time, a mixture of four parts of iron filings and $2\frac{1}{2}$ parts of flowers of sulphur. The moment the mixture is thrown in, the cover should be placed on the crucible, to prevent, as far as possible the access of air. The black mass so obtained is very like cast-iron. It may be broken from the crucible, when cold, with a chisel. A compound of ferrous and ferric sulphides occurs native, and is called *magnetic pyrites*, $\text{Fe}_3\text{S}_4 = \text{FeS}, \text{Fe}_2\text{S}_3$.

If you moisten ferrous sulphide with water, and let it remain exposed to the air in a warm place, small green crystals will be found disseminated throughout the mass, the sulphur having gradually attracted oxygen from the air. FeS thus becomes FeSO_4 .

Iron disulphide.—*Iron pyrites*, FeS_2 .—Iron combined with twice as much sulphur as in the proto-sulphide, occurs native in many ores, and frequently in common coal, and is called *iron pyrites*. It has the appearance of brass, and usually occurs in cubic crystals. If heated in a retort, a portion of the sulphur distils over, and is collected, and magnetic pyrites remains behind. *Green vitriol* is prepared from the residue, by piling it in heaps, and leaving it for several months exposed to the air. The green vitriol thus formed is freed from earthy impurities by lixiviation and evaporation. When iron pyrites is heated in the air, both its elements are oxidized:



The mineral is therefore often used instead of sulphur in the manufacture of sulphuric acid.

Tests for Iron.—The best test for iron is potassium ferro-cyanide (p. 187), which gives, with ferrous salts, a white or pale blue, and with ferric salts, a dark blue precipitate.

CHROMIUM.

Cr = 52.

Chromium in the pure state is little known; from the difficulty attending its isolation it can only be obtained in small quantities. It is remarkable for its hardness, and for its infusibility, which even exceeds that of platinum.

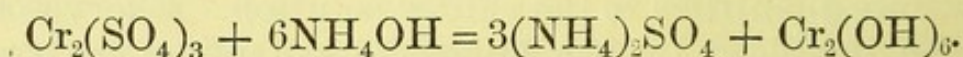
The chief ore of chromium contains chromic oxide and ferrous oxide ($\text{FeO}, \text{Cr}_2\text{O}_3$), and is called *chrome-iron-stone*, or *chrome-iron-ore*. Chromium likewise occurs in a mineral crocoisite, which is a lead salt of chromic acid, PbCrO_4 , and in small quantities in many other minerals, but its distribution is by no means abundant.

Experiment 1.—Grind a few grains of chrome-iron-ore to fine powder, and mix them with equal quantities of powdered nitre and potassium carbonate. Place the mixture in a small iron spoon and heat strongly with the blow-pipe or in a fire for several minutes. When the ignited mass is cold, detach it from the spoon and boil with a little water in a test-tube, and filter the solution. It will be of a bright yellow colour from the presence of a salt called *potassium chromate*, K_2CrO_4 ; a compound which will be described further on. Add dilute sulphuric acid in minute quantities to the yellow solution. Effervescence will be produced by the unaltered potassium carbonate. The addition of acid is stopped the instant the colour of the liquid changes to orange-red. The acid removes half the potassium, and a new salt, called *potassium dichromate*, $\text{K}_2\text{Cr}_2\text{O}_7$, is formed. This salt may be obtained by the evaporation of its solution in beautiful prismatic orange-red crystals. Immense quantities of potassium dichromate are prepared from chrome-iron-ore by a method very similar to the one just described. From this salt, directly or indirectly, all other chromium compounds are formed.

The constitution of the chromium compounds has already been indicated. The *chromous* compounds are difficult to prepare, and are of no interest. Of the *chromic* compounds the following are the most important:

Chromic oxide, Cr_2O_3 .—*Experiment 2*.—Mix a few grains of powdered potassium dichromate with a quarter of its weight of starch, and ignite the mixture in the iron spoon. After the ignition, boil the mass with some water. This will dissolve out the potassium carbonate formed during the reaction, and will leave chromic oxide as a green powder. Wash the powder once or twice, and dry it. This substance is employed as a pigment, and is valued for its permanence. It likewise produces a fine green on glass and porcelain, and is accordingly esteemed as a valuable vitrifiable pigment. It may be prepared in many ways, and of various shades of brilliancy. It is obtained in a beautiful form by strongly heating potassium dichromate alone.

Chromic hydrate, $\text{Cr}_2(\text{OH})_6$.—*Experiment 3*.—Mix some sulphuric acid with a solution of potassium bichromate, heat the mixture and gradually add methylated spirit. The brown colour of the solution will soon be changed to bright green. The sulphuric acid, as we shall afterwards find, decomposes the bichromate and sets free chromic acid. This is in turn decomposed by the spirit and reduced to chromic oxide, which is instantly dissolved by the excess of sulphuric acid present to form chromic sulphate, $\text{Cr}_2(\text{SO}_4)_3$, to which the green colour is due. Add ammonia in excess to the green solution, a bulky green precipitate is produced. It is chromic hydrate:



From chromic hydrate any of the per-salts of chromium may be prepared by solution in acids; thus, with hydrochloric acid, chromic chloride is obtained, with sulphuric acid chromic sulphate, &c. These salts are green. A corresponding series of violet chromic salts, of similar constitution to the green, is likewise known.

Chromic anhydride, CrO_3 .—*Experiment 4*.—Add powdered potassium dichromate to water until it ceases to dissolve, and filter or pour off the clear saturated solution so

obtained. Measure four drachms of the solution, and add to it gradually five drachms of strong sulphuric acid, and allow the mixture to stand. When the solution is cold, a mass of red needle-like crystals will be found at the bottom of the vessel. Pour off the solution and scrape out the crystals with a glass rod on a new brick or tile, cover with a tumbler and leave them to drain. These crystals consist of chromic anhydride. Chromic anhydride is a powerful oxidising agent; its action on combustible bodies may be exemplified by the following experiments:

Experiment 5.—Rinse out a tumbler with strong alcohol, then throw into it a few grains of chromic anhydride; the alcohol which remains adhering to the tumbler will combine with half the oxygen of the chromic anhydride, with such energy, that it ignites and instantaneously bursts into flame. The respective odours of acetic acid and aldehyde may be detected as the products of the oxidation of the alcohol, while the chromic anhydride is reduced to chromic oxide.

Experiment 6.—Mix in a small mortar as much chromic anhydride as can be taken up on the point of a knife with about one quarter as much of powdered camphor (without pressing upon it strongly), and then let some drops of alcohol fall from a considerable height into the mortar; instantaneous ignition and deflagration ensue, almost as if you were burning gunpowder. The residue in the mortar presents, after the decomposition, the appearance of a beautiful green mossy vegetation; it consists of chromic oxide, which at the moment of its formation was scattered by the burning camphor fumes, and was thereby most delicately subdivided.

Chromic anhydride dissolves in water to form a solution of a strongly acid character, and it is hence assumed to be a solution of chromic acid, H_2CrO_4 .

Potassium chromate, K_2CrO_4 .—*Experiment 7.*—Add potassium carbonate to solution of chromic acid until effervescence ceases: potassium chromate is thereby formed, as shown by the change of colour, from brown to yellow, of the solution. By evaporation the salt may be obtained in lemon-yellow crystals.

Experiment 8.—Add a solution of potassium chromate to

solutions of barium nitrate, lead acetate and silver nitrate, or to any other soluble salts of these metals. With the two former, light yellow precipitates of the chromates of the respective metals, with the latter a dark red one, will be produced. The lead chromate is used as a pigment, under the name of *chrome yellow*. If chrome yellow is digested with a cold solution of caustic soda, *chrome red* $\text{PbCrO}_4, \text{PbO}$, is obtained.

Besides the chromates, a second series of salts, called *dichromates*, is known. The potassium member of this series we are already familiar with. It is by far the most important of the chromium compounds. The process employed for its preparation has already been given, and during the preceding experiments many of its properties will have been examined. The formula for potassium dichromate is $\text{K}_2\text{Cr}_2\text{O}_7$. It may be regarded as a compound of potassium chromate with chromic anhydride ($\text{K}_2\text{CrO}_4, \text{CrO}_3$) and is therefore sometimes called an *anhydro-chromate*. The rest of the dichromates are of analogous constitution. As a rule, the dichromates are orange-red, while the chromates are light yellow.

Ammonium dichromate, $(\text{NH}_4)_2\text{Cr}_2\text{O}_7$.—*Experiment 9*.—This compound is interesting from the curious way it is decomposed on being heated. Divide a solution of chromic acid into two equal parts, and cautiously add ammonia to one of them until the brown colour changes to yellow. If now the solution were evaporated, yellow crystals of neutral ammonium chromate would be obtained, but on adding the second part of the chromic acid, and setting aside in a warm place to evaporate, red crystals of ammonium dichromate will be formed. Collect some of the crystals, dry in a fold of filtering paper, and heat them in a dry test-tube or on a piece of tinned-plate. The salt will decompose energetically, pure nitrogen will be evolved, and the crystals will lose their shape and colour, and be converted into an irregular green mass of chromic oxide, bearing some resemblance to a heap of green tea.

Experiment 10.—Another pretty experiment, founded on the same decomposition, may be made by soaking filtering-paper first in tincture of benzoin, and then in a

strong solution of ammonium dichromate, and drying it. On folding the paper in the shape of a fan, and applying a light to one corner, it gradually smoulders away, and a dark green residue is left, having the appearance of fern leaves.

Chloro-chromic anhydride, CrO_2Cl_2 .—*Experiment 11.*—This interesting compound may be prepared by mixing one ounce of powdered potassium dichromate with an equal amount of common salt, putting the mixture in an earthen crucible, and heating strongly in a fire. The contents will soon melt, and must then be poured on a flat stone or metal surface to solidify. The hard cake so obtained is broken into small pieces, and introduced with three ounces of strong sulphuric acid into a retort of moderate size. On applying heat, brown vapours soon fill the vessel and condense to a dark-red liquid, which may be received in a flask or bottle. This liquid is chloro-chromic anhydride, which may be regarded as chromic anhydride, CrO_3 , having one atom of oxygen replaced by two of chlorine. It is not, however, miscible with, or soluble in water, but in contact with that liquid is soon decomposed into chromic acid and hydrochloric acid. Its action on combustible substances resembles that of chromic anhydride, but is more violent. This may be shown by pouring a drop of the liquid on a little alcohol, benzol, or turpentine. In either case, the action which ensues is sufficiently intense to inflame the mixture. A minute fragment of phosphorus thrown on a drop of the liquid produces a violent explosion. By the same substance many other combustible bodies, sulphur, sugar, &c., may be inflamed by mere contact.

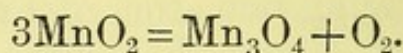
MANGANESE.

Mn = 54.

Like chromium and many other metals, manganese is so difficult to separate from its ores that it has not hitherto been applied to any useful purpose, although some of its compounds are of great importance. Manganese is an exceedingly hard, brittle, reddish-white metal, which rapidly oxidises when exposed to the air, and decomposes water slowly, even at ordinary temperatures. Several

compounds of manganese occur native, chiefly oxides, the principal one being black oxide or manganese peroxide, MnO_2 , which is found in great abundance as the mineral *pyrolusite*, from which the metal may be obtained by intense ignition with charcoal. The compounds of manganese are analogous to those of chromium, but, unlike the latter, the *manganous* compounds are more stable than the *manganic*.

Oxides of Manganese.—Several oxides of manganese are known. *Manganous oxide*, MnO , *manganic oxide*, Mn_2O_3 , both unimportant; a red oxide, Mn_3O_4 , analogous to the magnetic oxide of iron, which is met with as the mineral *hausmannite*; and finally, the most important of all manganese compounds, the *peroxide*, *binoxide*, or *black oxide* as it is often called MnO_2 . Many of the properties of this oxide are already familiar to the student. When heated to redness it loses part of its oxygen, and the red oxide remains:



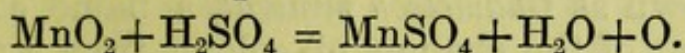
This fact, as well as its use in the preparation of chlorine, iodine, and bromine, has already been treated of in the description of those substances.

Glass-makers often add manganese binoxide to green bottle glass, which contains iron in the ferrous condition, to change its colour to brown, which it effects by converting the iron from the ferrous to the ferric state. When fused with white glass it communicates a violet colour to it. In this way imitation amethysts are made. From manganese binoxide the other compounds of the metal may be obtained, either directly or indirectly.

Manganous sulphate, MnSO_4 .—*Experiment 1.*—Mix in a porcelain crucible a quarter of an ounce of manganese peroxide with one-eighth of an ounce of sulphuric acid, and expose the mixture to a gentle heat for fifteen minutes, and then to a strong heat for an hour. After cooling, boil the black mass in water, filter and evaporate the solution to dryness, constantly stirring it when nearly dry; the reddish-white powder is manganous sulphate, generally accompanied with iron derived from the ore.

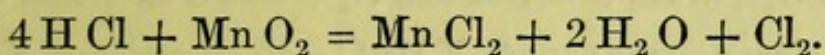
The effervescence with which the action of the sulphate

acid is attended is due to the escape of oxygen, which might easily be collected if the experiment were performed in a suitable apparatus. This explains why a mixture of sulphuric acid and manganese peroxide is frequently used for oxidising purposes. The action which occurs is thus represented :



Manganous chloride, MnCl₂.—*Experiment 2.*—Heat strong hydrochloric acid with pure manganese peroxide for half an hour, filter the solution and evaporate to a small bulk. On cooling, pink crystals of manganous chloride will be obtained.

During the heating effervescence will occur, owing, in this case, to the escape of chlorine. In fact, it will be remembered, chlorine is most easily prepared by these very materials, and accordingly manganous chloride is obtained as a bye-product when they are employed :



Dissolve a portion of manganous sulphate or chloride free from iron in water, and use it for the following experiments :

Experiment 3.—Add ammonia or potash to a portion of the solution. A white precipitate of manganous hydrate will be thrown down. It rapidly turns brown and becomes manganic hydrate.

Experiment 4.—Add ammonium sulphide to another portion of the solution. A buff-coloured precipitate of manganous sulphide is produced.

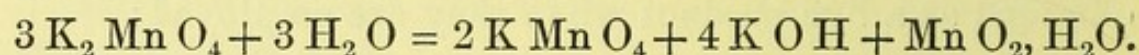
Manganic and Permanganic Acids.—Two very interesting series of salts are known, the acid radicals of which are derived from oxides of manganese, one of which is known. They are called the *manganates* and *permanganates*. The potassium salts are the most important :

	Anhydride.	Acid.	Potassium Salt.
Manganic	Mn O ₃ ?	H ₂ Mn O ₄ ?	K ₂ Mn O ₄ .
Permanganic	Mn ₂ O ₇	H Mn O ₄	K Mn O ₄ .

Potassium manganate, K₂MnO₄.—*Experiment 5.*—Mix intimately in a mortar one drachm of potassium carbonate, one drachm of manganese peroxide, and half a drachm of

potassium nitrate; put the mixture in a crucible and heat strongly in a fire for half an hour. When cold add a little water to a small portion of the mass, and allow it to stand and settle. A dark-green solution will be obtained of potassium manganate, analogous to potassium chromate, K_2CrO_4 , which is similarly obtained. Potassium manganate is known as *Chameleon mineral*, a name conferred on it on account of the change of colour from green to red that it undergoes on its solution being diluted with water.

Experiment 6.—Pour a little of the potassium manganate solution into a tumbler of water. The dark-green colour of the salt instantly begins to change, passing through various shades of green and violet until it becomes crimson-red. This remarkable change is owing to the conversion of the salt into permanganate and is effected more easily by dilute acids. A hydrate of manganese peroxide is formed at the same time:



Potassium permanganate, $KMnO_4$, often written $K_2Mn_2O_8$.—This salt, like the manganate, is remarkable for the facility with which it parts with its oxygen to oxidisable substances.

Experiment 7.—To some of the potassium permanganate prepared in the previous experiment add a few drops of sulphurous acid. The red solution instantly becomes colourless. The sulphurous acid takes the oxygen from the permanganate, and a colourless manganous salt is formed.

Perform the same experiment with the green potassium manganate, a like result will ensue. Even a piece of wood, paper, or any other organic substance, thrown into the green or red solutions, decomposes and removes their colour, and for this reason they should not be filtered through paper.

On account of the ready oxidation of organic matter effected by these salts, they are often used as disinfectants, and they are now prepared in large quantities and sold as "Condyl's fluids," from the name of the manufacturer. They rapidly oxidise and destroy noxious gases or putrefying substances, and indeed, almost any kind of organic matter.

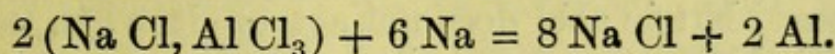
Experiment 8.—The unwholesome gas, hydrogen sulphide, is often evolved from decomposing animal matters. Mix a little of the green or red salt with a solution of hydrogen sulphide. The bad odour of the gas is instantly destroyed.

Experiment 9.—Procure some water known to be of bad quality, as from a stagnant ditch, acidulate it with sulphuric acid, and add some permanganate to it gradually. A large quantity of the salt will be decolourized. Repeat the experiment with good drinking water; the colour of permanganate will remain, even though very little be added. In fact, a sort of rough estimation of the amount of impurity present may be made by observing the relative quantities of permanganate which different waters decolourize.

ALUMINIUM.

Al = 27.

Aluminium, which derives its name from one of its most important compounds—alum, exists in greater abundance than any other metal. It is only separated from the compounds which contain it with great difficulty, and at one time was only obtainable in the smallest quantities. New processes, however, have materially facilitated its extraction, and it is now applied to many useful purposes. It occurs in numerous minerals of complex constitution, in shales, slates, clays, and in all other older rocks. As oxide it forms the valuable minerals corundum or emery, and the gems sapphire and ruby, the two latter being coloured respectively, with cobalt and chromium. A great number of silicates containing aluminium are known, of which felspar, one of the constituents of granite, is among the most important. Aluminium is most easily obtained by the action of sodium, at a high temperature, on a peculiar compound of the chlorides of aluminium and sodium:



A double sodium and aluminium fluoride, 3NaF, AlF_3 , found native, and known to mineralogists as *cryolite*, is sometimes used instead of or with the double chloride.

Aluminium is a bluish white metal, somewhat resembling zinc in appearance. It does not tarnish on exposure to air, and from this property, and the fact that it is only about one-fourth the weight of silver (Sp. Gr. 2.56), it is applied to many useful and ornamental purposes. The beautiful alloy called "aluminium gold" is an alloy of one part of aluminium with nine parts of copper.

COMPOUNDS OF ALUMINIUM.

Aluminium oxide, Al_2O_3 , *Alumina*.—This substance, it has already been mentioned, is found native as corundum, &c. When pure it is white. It may be formed artificially by heating the hydrate. At a very high temperature rubies and sapphires, identical with those found in Nature, have been prepared.

Aluminium hydrate, $\text{Al}_2(\text{OH})_6$.—*Experiment 1*.—Dissolve a little alum in water, and add sodium carbonate in excess; a bulky white precipitate of aluminium hydrate will be thrown down. Carbonic anhydride escapes as gas, for aluminium carbonate does not appear to exist.

Experiment 2.—Add potassium hydrate to a little of the white precipitate; it instantly dissolves. *Potassium aluminate*, $\text{K}_2\text{Al}_2\text{O}_4$, is formed. Aluminic hydrate plays in this case the part of an acid.

Experiment 3.—Add potassium hydrate cautiously to the alum solution; aluminium hydrate is precipitated, but on adding excess is again dissolved, as in the last experiment. If ammonia is used instead of potash, the same precipitate is formed, but does not dissolve in excess.

Experiment 4.—Put a piece of thin sheet aluminium into potassium hydrate solution. It dissolves slowly with effervescence, but more rapidly if heated. The effervescence is due to the escape of hydrogen. Potassium aluminate is formed. Aluminium dissolves in hydrochloric, and in dilute sulphuric acids, with production respectively of chloride and sulphate, and escape of hydrogen. Nitric acid does not act on it when cold, and only with difficulty when boiled with it.

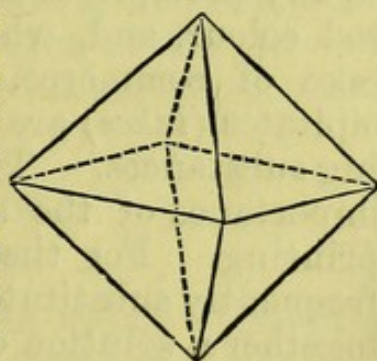
Aluminium sulphate, $\text{Al}_2(\text{SO}_4)_3$.—*Experiment 5*.—Dry thoroughly a piece of white clay, and expose it for some hours to heat, which is most easily done on the plate of a

heated oven; then rub two ounces of it to a powder in a porcelain bowl with one ounce of sulphuric acid; pour upon the mixture one ounce of water, and let it remain some weeks in a warm place. Frequently stir the mass during this time with a glass rod. Finally, dilute it with six ounces of boiling water, and strain it through linen. The residue on the latter consists principally of silica, but aluminium sulphate is found dissolved in the liquid.

Clay is an insoluble aluminium silicate, which, on being treated with sulphuric acid in the manner described, is decomposed, aluminium sulphate being formed and silica set free. Evaporate the liquid so obtained till only one and a half or two ounces of it remain, and then put it in a cool place; it will crystallize in thin silky plates of a pearly lustre, which are very deliquescent. Pour off the liquor remaining behind, which always contains free sulphuric acid, and again dissolve the crystals in a little water. In factories the solution is frequently evaporated to dryness, and a solid mass is thereby obtained, which is employed in calico-printing and dyeing.

Alum, $K_2Al_2(SO_4)_4$.—*Experiment 6*.—Saturate two ounces of boiling water with potassium sulphate, and add to it the solution of aluminium sulphate obtained in the last experiment. Stir the mixture till it is cold, and decant the clear liquor from the white sediment. The sediment is *alum* in a state of powder. If dissolved in boiling water and *slowly* cooled, you obtain from it crystallized alum in beautiful, transparent, four-sided double pyramids (octahedrons).

Fig. 86.



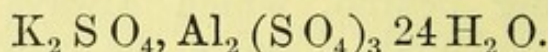
Alum is a double potassium and aluminium sulphate, $K_2SO_4, Al_2(SO_4)_3$, and is one of the most important of aluminium compounds. It is obtained from various sources; in the country it is prepared from a kind of shale, called *alum shale*, or *alum ore*, an aluminous mineral, which contains an abundance of iron pyrites, FeS_2 . When the shale is slowly roasted, the pyrites becomes ferrous sulphate, $FeSO_4$, while the second atom of sulphur takes

oxygen, and becomes sulphuric acid, which decomposes the aluminium silicate and forms aluminium sulphate. On the addition of potassium sulphate, alum is formed, and may be purified by recrystallization.

At Tolfa, in Italy, native alum is found, in conjunction with aluminium hydrate, called alum-stone. From this it is prepared by gently heating to decompose the hydrate, and then extracting the alum with water. It is sold as Roman alum, and is valued on account of its freedom from iron, which always contaminates ordinary kinds.

Experiment 7.—Heat a small crystal of alum before the blow-pipe; it foams and melts, and ultimately yields a white porous mass (burnt alum). The foaming is owing to the evaporation of the water of crystallization, which constitutes nearly one-half of the weight of the alum.

The formula for crystallized alum is



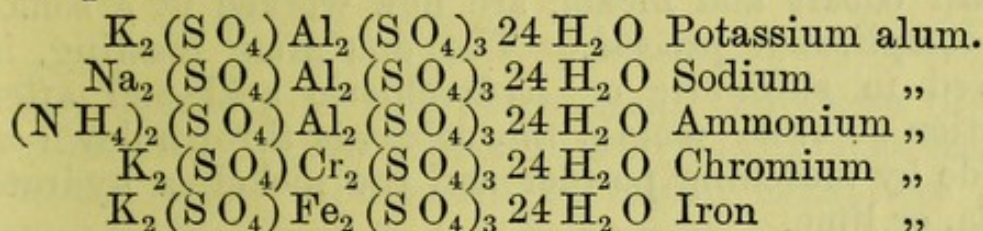
Experiment 8.—Boil half an ounce of Brazil-wood for fifteen minutes in six ounces of water, decant the decoction, and dissolve in it half an ounce of alum; it thereby acquires a more brilliant red colour. Now add to it a solution of sodium or potassium carbonate, as long as any precipitate is produced; this precipitate is of a fine red colour, and, when dried, constitutes the Brazil-wood lake of commerce. In a similar manner, coloured precipitates (*lakes*) are obtained from other vegetable colouring substances. This example accounts for the great importance of the aluminium salts in dyeing and calico-printing. For these purposes aluminium acetate is very frequently substituted for alum. It is obtained by mixing together a solution of lead acetate and aluminium sulphate (or alum), whereby *soluble aluminium acetate* (alum mordant) and *insoluble* lead sulphate are formed.

Experiment 9.—Moisten a very small piece of alum with a drop of a solution of cobalt nitrate, and heat it before the blow-pipe; the oxide of cobalt, which remains behind, imparts a *beautiful blue* colour to the compound of aluminium. This fact is frequently taken advantage of, as a very accurate means of detecting aluminium.

Another splendid blue pigment, artificial *ultramarine*, can

be prepared by heating to redness a mixture of white clay, sand, sulphur and resin. Its nature is not known. This pigment must be carefully kept from contact with acids, as they would evolve from it sulphuretted hydrogen, and destroy the colour.

Alum affords a fine example of *isomorphism* (p. 84). We are able to replace the potassium in the alum by another monad metal, as sodium or ammonium; or to replace the aluminium by a similar metal, as chromium or iron, without thereby changing the octahedral crystalline form. We thus obtain alums of which the following are examples:



The first three of the above alums have a white colour; chromium alum, a deep red; and iron alum, a pale violet colour. They may be prepared by dissolving together in water their constituent salts, in proper proportions, and allowing the solution to crystallize.

Isomorphous compounds cannot be separated by crystallization. If a crystal of common alum is hung in a cold saturated solution of iron alum, a layer of the new salt will form over the first, and this process may be repeated with other alums.

COBALT AND NICKEL.

Co = 59. Ni = 58.

During the Middle Ages, when the miner believed in the existence of earth-spirits and goblins in the solitary depths of his mines, ores were occasionally found, particularly in the mines of Schneeberg, in Saxony, resembling in brilliancy and weight the finest silver ores, which, however, yielded in the smelting furnaces no silver, but crumbled away to a grey ash, a disagreeable odour of garlic being at the same time emitted. In accordance with the superstitious notions of those times, the miner attributed the disappearance of the supposed

silver to the malicious jests of the earth-spirits, and contemptuously rejected these ores, which he baptized by the names—cobalt and nickel, which they still bear. But now they are held in high estimation, cobalt being used for imparting a beautiful blue colour to glass and porcelain, and nickel for giving to brass the appearance of silver. As these metals are melted only with great difficulty, the heat of the old furnaces was not sufficient to fuse them. The odour of garlic was occasioned by the arsenic, which always accompanies the ores of cobalt and nickel.

The ores (speiss-cobalt, cobalt glance, &c.), containing arsenical cobalt and nickel, are now worked by a somewhat complicated process. The ore, after roasting, is dissolved in sulphuric or hydrochloric acid and after separation of other metals the cobalt is thrown down as peroxide by bleaching powder and the nickel as hydrate by soda, or lime.

Smalt is a potash glass coloured blue by cobalt silicate.

As pure metals, cobalt and nickel have a great similarity to iron, both in their external appearance and in their combinations; but they do not attract oxygen with such avidity. The three metals, iron, cobalt, and nickel, constitute, as has been already mentioned, the *magnetic trio*; they alone, of all the metals, are attracted by the magnet to any appreciable extent. It is, moreover, remarkable, that just these three metals always occur in meteorites, which occasionally fall to the earth, we know not whence, in a glowing state (meteoric iron, meteoric stones). Nickel is largely used for plating iron (nickel plating). The operation is conducted as with copper (p. 227), except that the solution used is nickel-ammonium sulphate, $\text{NiSO}_4(\text{NH}_4)_2\text{SO}_4$, and the positive pole is a sheet of pure nickel.

Cobalt oxide, CoO , is of an olive-green colour, and its *hydrate* is pink; *cobalt peroxide*, Co_2O_3 , is black. The oxide is frequently employed in staining glass.

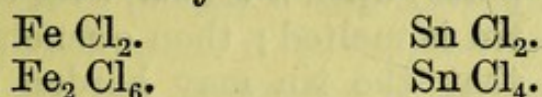
Nickel oxide, NiO , is of a greenish-grey, and its *hydrate* of a beautiful apple-green colour; the *peroxide*, Ni_2O_3 , is black. *Chrysoprase*, known as an ornamental stone, is quartz, coloured green by nickel.

The *proto-salts of cobalt* are of a pink colour. A solution of *cobalt nitrate* is often used in blow-pipe experiments; a solution of *cobalt chloride* is employed as a sympathetic ink, as it possesses the property of becoming blue when the water evaporates, and again pink on absorbing water. Cobalt forms, with phosphoric and arsenious acids, red insoluble compounds, which are now employed as vitrifiable pigments in glass and porcelain painting. The *salts of nickel* have a light-green colour.

The salts of cobalt and nickel, like those of iron, are not precipitated in acid solutions by H_2S , but are thrown down as black sulphides by ammonium sulphide.

GROUP ii.—Tin, Platinum.

The members of this group form two series of compounds, a *diad* series, represented by $SnCl_2$, SnO , &c., and a *tetrad* series, represented by $SnCl_4$, SnO_2 , &c. The former is analogous to the diad series, but the latter is not analogous to the tetrad series in Group i. In fact, the metals of the iron group do not yield simple tetra-compounds, but only compounds in which two atoms of metal act as a single hexad atom. The formulæ for the chlorides of iron and tin sufficiently illustrate the difference:



TIN.

$$Sn = 118.$$

Tin is one of the few metals which were known in the most ancient times. It becomes fluid at a very moderate heat, namely, at $446^\circ F.$ ($230^\circ C.$), and in many countries its ores are found in the sand with which the surface of the soil is covered; therefore it is easily obtained and easily smelted. Formerly it was obtained principally from the British Islands, which were, therefore, called the tin islands, and even at the present time they, together with Malacca in the East Indies, furnish the purest tin.

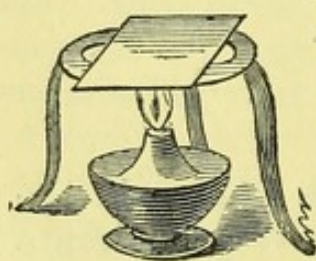
Tin almost exclusively occurs as an oxide, SnO_2 , called *tin-stone*. The distribution of this ore in England is chiefly confined to Cornwall. Large quantities are also obtained from the Malay Peninsula, Australia and elsewhere.

Tin is prepared in a very simple manner from tin-stone. The finely ground and washed ore is first roasted, by which process the arsenic and sulphur are volatilized and the iron oxidized. The iron and copper are then in a great measure removed by washing. Finally, the ore is reduced with charcoal or anthracite (highly carbonaceous coal).

The properties which especially characterize tin, and render it a very valuable metal, are its beautiful lustre, and its great softness and flexibility,—its slight affinity for oxygen, in consequence of which it long retains its brightness in the air and in water,—its easy fusibility—it melts at 235° C. (455° F.),—which renders it peculiarly well adapted for casting, and for coating other metals (*tinning*). Most of the tin of commerce contains traces of arsenic and other metals. A bar of tin emits a grating sound on being bent, and by repeating the operation several times in succession it becomes very hot; the tin, on hardening, assuming a crystalline texture. These crystals may be very beautifully produced upon tinned-iron sheets.

Experiment 1.—Heat a piece of tin-plate (tinned-iron plate) upon a tripod, over a lamp, till the tin is melted; then quench it with water,

Fig. 87.



that the tin may harden quickly. The surface of the plate has a dull grey aspect, for it is covered with a film of oxide; but the most beautiful crystalline figures will very soon appear upon it by rubbing it alternately with balls of paper, one of which is moistened with diluted aqua regia, and the other with potash. Both these liquids dissolve the coating of oxide, and lay bare the pure metallic tin surface (*moiré métallique*).

Alloys of tin and lead are generally used by workers in metal, under the name of *solder*, for joining metals together (soft soldering). Solder is to the tinman what glue is to the carpenter. An alloy of two parts of tin and one part of lead is the most easily fusible, and is called *fine solder*. Another alloy, used in the soldering of coarser articles, such as gutters, is composed of two parts

of lead and one part of tin, and is called *coarse solder*; it is so thick that it does not spread itself, but must be applied by smearing. For soldering those metallic articles which are to be subjected to a stronger heat, brass, or some other alloy of difficult fusibility, is made use of (*hard solder or brazing*).

Some lead is added even to the tin of which the tinman makes his articles, because pure tin is somewhat brittle, and does not adapt itself well to the moulds. The quantity of lead which can be added to tin is in many countries regulated by law ($\frac{1}{9}$ to $\frac{1}{6}$). Such an alloy is called *proof tin*, to distinguish it from refined or grain tin, which is tin in its greatest purity.

Gun-metal contains copper 9 and tin 1; *Speculum-metal*, copper 2, tin 1; *Bell-metal*, copper 5, tin 1; *Bronze* varies much, consisting generally of copper, tin, and lead, or zinc.

Tin is very malleable: sheets of tin-foil may be obtained not more than $\frac{1}{1000}$ th of an inch thick. Spurious *silver-leaf* is an alloy of tin and zinc hammered out into extremely thin leaves.

An alloy of tin and mercury, or *tin-amalgam*, forms the silvering of looking-glasses. Tin does not readily tarnish in air, and on this account is frequently employed as a coating for the more oxidisable metals copper and iron. For coating the latter, as in the manufacture of tin-plate, clean sheets of the best iron are first immersed in melted fat, and then in melted tin for some time, and drained. For coating copper, melted tin is poured on the clean surface of the metal, and rubbed over it with a wisp of tow.

Experiment 2.—The tinning of brass or copper may also be done in what is called the moist way, by heating to the boiling point finely cut up tin-foil or tin scrapings in a pot with cream of tartar and water, and then boiling in this liquid for half an hour brightly polished copper or brass articles; as, for instance, farthings and brass nails. In this manner, pins, which consist of brass wire, can be tinned or whitened.

The symbol Sn for tin is derived from its Latin name, *stannum*.

COMPOUNDS OF TIN.

Stannous oxide, SnO .—*Experiment 1*.—Pour some ammonia upon a solution of stannous chloride; the white precipitate which is formed is *stannous hydrate*. By boiling the liquid, the hydrate is decomposed into water and stannous oxide, which has a black, or dark-green colour, and must be quickly washed with boiled water and dried, as it easily attracts more oxygen from the air. If you heat the dried oxide before the blow-pipe, it burns with great briskness, like tinder, forming stannic oxide.

Stannic oxide, or *Stannic anhydride*, SnO_2 .—*Experiment 2*.—Heat a piece of tin upon charcoal in the outer blow-pipe flame; it will soon become covered with a powder, of a yellow colour when hot, but white when cold; this is stannic oxide. Tin peroxide thus obtained is not soluble in any acid, and can only be fused with great difficulty. It is so fine a powder, that it is often used for polishing glass and metals. For this purpose it is sold as *putty-powder*.

Experiment 3.—Mix a grain or so of stannic oxide with a few grains of charcoal, and heat the mixture on a piece of charcoal with the blow-pipe. The oxide will be reduced and a bead of metallic tin will be formed. This experiment serves to illustrate the process used for the extraction of tin.

Stannic oxide resembles silica in its structure, and in its relation to water. Although insoluble in water, it is a true anhydride and two distinct acids, known respectively as *stannic* and *metastannic* acids, can be prepared by indirect means. They are both compounds of SnO_2 with water.

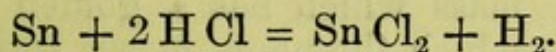
Stannic acid, H_2SnO_3 .—*Experiment 4*.—Add ammonia in excess to solution of stannic chloride. A white precipitate falls, which, when dried in the air, has the formula H_2SnO_3 . It dissolves in potash and soda, forming salts called *stannates*. Sodium stannate is largely used as a mordant in calico-printing.

Metastannic acid.—*Experiment 5*.—Metallic tin is rapidly oxidised by strong nitric acid; red fumes are given off and a white powder remains. This is *metastannic acid*.

Dried in the air, its composition corresponds with the formula H_4SnO_4 , but when heated to 100°C . (212°F .), it loses water, and its composition is then represented by the formula H_2SnO_3 . The *metastannates* are ill-defined salts, and their constitution is still uncertain.

Both the above acids are converted by heat into stannic anhydride and water.

Stannous chloride, SnCl_2 .—*Experiment 6*.—Boil some fragments of pure granulated tin (the tin of commerce often contains lead) with strong hydrochloric acid. Hydrogen escapes, and stannous chloride is formed, and may be obtained in colourless crystals ($\text{SnCl}_2 \cdot 2\text{H}_2\text{O}$) by evaporation :



The solution of stannous chloride on exposure to air becomes milky, and changes in part to stannic chloride.

Stannic chloride, SnCl_4 .—*Experiment 7*.—Add chlorine water to a solution of stannous chloride, until the odour of chlorine is no longer destroyed. SnCl_2 is thereby converted into SnCl_4 . This combination can also be obtained by dissolving tin in aqua regia (nitric and hydrochloric acids). The dyers call this liquid tin mordant, or red spirits. By the addition of ammonia, stannic acid is obtained.

When pure, stannic chloride is a volatile liquid.

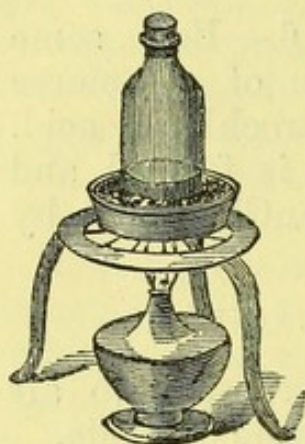
Stannous sulphide, SnS .—*Experiment 8*.—Add sulphuretted hydrogen to an acid solution of stannous chloride. A reddish-brown precipitate of stannous sulphide is produced. It may also be obtained by the direct combination of tin and sulphur.

Experiment 9.—Envelop twelve grains of flowers of sulphur in a piece of tin-foil, weighing twenty-four grains, then roll up the packet so that it may be introduced into a test-tube, and heat it; half of the sulphur burns, but the other half combines with the tin, forming a brownish-black mass of a metallic lustre (SnS). If you sprinkle the tube, while still hot, with water, it is rendered friable, and can easily be separated from the fused sulphide. The weight of the latter amounts to nearly thirty grains.

Stannic sulphide, SnS_2 .—*Experiment 10*.—Add sulphu-

retted hydrogen to solution of stannic chloride. In this case a bright *yellow* precipitate of stannic sulphide is produced. It may be prepared of a beautiful metallic appearance as follows:

Experiment 11.—Pulverise thirty grains of stannous sulphide obtained by a former experiment, and mix the powder intimately with six grains of sulphur and twelve grains of sal-ammoniac; put the mixture into a thin-bottomed glass flask of an ounce capacity, and heat it strongly for an hour and a half in a sand-bath. You obtain stannic sulphide; but in this case, as a mass which has a golden lustre, and to which the name *aurum musivum*, or *mosaic gold*, has been given. It may be used for giving a gold-like coating to wood, plaster of Paris, clay, &c. (*bronzing*). The sal-



ammoniac is found again as a sublimate in the upper portion of the flask; it prevents by its volatilization the mixture getting hot enough to destroy the colour of the stannic sulphide.

PLATINUM.

Pt = 195.

Platinum, one of the heaviest of the elements (Sp.Gr. 21.5), was brought in the last century from America, where it was found, in the form of small flattened grains, mixed with the sand from which the gold was washed. It received the name of *platinum*, derived from the Spanish word *plata*, silver, on account of its resemblance to silver in colour and ductility. It was afterwards found also in the sand of the Ural Mountains, in compact *lumps*, from the size of a lentil to that of a man's fist. Its occurrence in large masses is, however, very rare. The native platinum is not pure, but contains several other rarer metals in small quantities, namely, palladium, rhodium, ruthenium, iridium, and osmium. These metals can only be removed by a difficult process. Platinum, like gold, is a *noble* metal, and, like iron, is *tenacious*, *ductile*, and *can be welded*, and is, moreover, infusible at the strongest furnace

heat. These properties have rendered platinum an invaluable metal to the chemist. Sulphuric and hydrofluoric acids can be distilled in platinum retorts, nitric acid can be boiled in platinum capsules, and substances can be subjected to the strongest white heat in platinum crucibles, or on platinum foil or wire, without the platinum articles being injured or melted. Great difficulties were at first experienced in working platinum. It was for a long time effected in the following manner: the platinum was first converted into a double chloride of platinum and ammonium by dissolving it in a mixture of nitric and hydrochloric acids, and then adding ammonium chloride. A yellow precipitate is thus produced, which on being heated strongly yields metallic platinum in a finely divided state (spongy platinum). The powder on being hammered in moulds, and then strongly heated and hammered again, becomes a compact block capable of being beaten into any shape. Although platinum is infusible by the heat from ordinary fuel, it can be melted at a temperature of $2,000^{\circ}\text{C.}$ ($3,632^{\circ}\text{F.}$) by the oxy-hydrogen blow-pipe, and at the present time furnaces made of quicklime are used, in which very large quantities of platinum can be fused by the aid of several jets of the burning gases. Platinum, like tin, forms a diad and a tetrad series of compounds.

Platinic chloride, PtCl_4 .—*Experiment 1.*—Heat some pieces of thin platinum foil or wire in a mixture of nitric and hydrochloric acids (*aqua regia*). The metal will gradually dissolve and form a brownish-yellow solution, which by evaporation on a water-bath yields a brown mass of platinum hydrogen chloride $\text{H}_2\text{PtCl}_6 = 2\text{HCl}, \text{PtCl}_4$. This is the chief salt of platinum, and from it the true chloride and other compounds are prepared. The above acid mixture forms the only available solvent of pure platinum, and neither will dissolve it separately. An alloy of platinum with much silver will however dissolve in nitric acid, platinic and silver *nitrates* being thereby formed. Platinum chloride combines with alkaline chlorides to form double salts on their solutions being mixed. The sodium salt is soluble, the rest are not.

Potassium-platinic chloride, $2\text{KCl}, \text{PtCl}_4$.—*Experiment 2.*—Mix solutions of platinic chloride and potassium chloride.

A yellow crystalline precipitate of the double chloride is produced, which is rendered more abundant by adding alcohol. Any other potassium salt may be employed if a few drops of hydrochloric acid be likewise added. The reaction thus affords a test for potassium.

Ammonium-platinic chloride, $2\text{NH}_4\text{Cl}, \text{PtCl}_4$.—*Experiment 3.*—Repeat the last experiment with solution of ammonium chloride. A yellow precipitate, similar in appearance to the preceding one, will be produced. It contains ammonium instead of potassium. Consequently, before testing for potassium it is necessary to prove the absence of ammonium.

Experiment 4.—Collect some of the yellow precipitate obtained in the last experiment, dry it in a basin and then heat it gently until it is quite grey. The compound will be decomposed, and metallic platinum only will be left as a grey, loosely coherent, porous mass—the so-called *spongy platinum*.

Spongy platinum, like charcoal, has a remarkable power of absorbing gases. On being introduced into a mixture of hydrogen and oxygen it absorbs those gases, and by thus bringing them into close contact causes them to unite, and an explosion is produced. A jet of hydrogen may be ignited in air by spongy platinum or platinum leaf by reason of this property. The apparatus for this experiment has already been described (page 108). Clean sheet platinum itself possesses this property, but in a less degree.

By proper chemical means, platinum may be divided still more minutely than in spongy platinum; it is then obtained in the form of a fine black powder, which possesses, in a still higher degree than spongy platinum, the power of condensing gases; it is called *platinum black*. If some alcohol be dropped upon this platinum black, ignition takes place, with an almost instantaneous oxidation of the alcohol. The reason of this change is to be sought for in a combination of the alcohol with the oxygen of the air, which is effected by means of the porous platinum black.

In many other cases oxidation may be promoted by platinum in these states of division.

CHAPTER IV.

METALLIC TRIADS, OR PENTADS.

GROUP i.—GOLD.

GROUP ii.—Arsenic, Antimony, Bismuth.

GOLD forms two chlorides, AuCl_3 and AuCl . It is, therefore, triad and monad. The metals of Group ii. are sometimes triad and sometimes pentad. They may, therefore, be classed with nitrogen and phosphorus; and indeed arsenic and antimony present so many points of analogy with those non-metallic elements that they are often excluded from the class of metals. This analogy is sufficiently indicated by the parallelism of the formulæ for the hydrides chlorides, oxides, and acids of the four elements. Bismuth also is triad and pentad, but, as it does not yield a hydride, its resemblance to the nitrogen group is less perfect than that of the other two.

GROUP i.—GOLD.

GOLD.

$$\text{Au} = 196.$$

Gold, like platinum, is always found in the metallic state. It is found in most countries, yet it is disseminated so sparingly, and the separation of it from the rocks or the river-sand in which traces of it occur is attended with so much labour, that it is rendered the most costly of our common metals. At present it is obtained most abundantly in Australia and California, and always contains silver, and frequently small quantities of other metals. It occurs in small grains, masses, or veins, in the older rocks and in their detritus, and in minute quantities in the sands of many rivers. To separate the gold from the sand or

detritus the latter is mixed or stirred up with water, when the gold, from its greater weight, sinks to the bottom, and the impurities which remain suspended are poured off with the water. Rocks containing gold are first finely crushed, and then shaken or mixed with mercury. The latter dissolves the gold. The mercury is then, in a great measure, removed by pressure in leathern bags, and finally and completely, from the resulting almost-solid gold amalgam, by distillation. In California the rocks are disintegrated by enormously powerful jets of water.

Gold is more malleable than any other metal, and is also very ductile. In making gold-leaf a single grain of the metal is hammered out until it covers the space of more than fifty square inches, or it may be drawn into a wire seven hundred feet long. The wire does not, however, possess much tenacity. It is far inferior to iron wire in this respect.

Gold-leaf is often so thin as to be translucent; that is, it transmits light without being transparent. This may be observed by carefully laying a gold-leaf on a plate of glass, and looking through it at a bright light. The film of gold transmits a green light. The colour becomes red if the lustre of the metal is destroyed by heat. Gold is not oxidised or tarnished by exposure to air or water, either at high or low temperatures, nor is it acted on by any common acid or alkali; consequently it is highly esteemed for ornamental purposes, and above all for coinage. Its specific gravity is 19.3, so that it is rather lighter than platinum. Its melting point is higher than that of silver but lower than that of copper.

Pure gold, like pure silver, is somewhat soft, and quickly wears out in using; therefore, when it is to be manufactured into coins or articles of luxury, it is alloyed with other metals, usually *silver* and *copper*, to render it harder. The quantity of pure gold contained in a mass is sometimes expressed by the word *carat*, the standard number not being 8, as in silver, but 24. A mark of gold (8 ounces) is divided into 24 parts, or carats. If gold is said to be 18 carats fine, it is understood that the mass consists of three-fourths (18 parts) of gold, and one-fourth (6 parts) of alloy; if 6 carats fine, of one-fourth (6 parts) of gold,

and three-fourths (18 parts) of alloy, &c. The gold used for coinage in England contains 22 parts gold to 2 of copper (22 carats).

Parting of gold.—In order to obtain fine gold from alloyed gold, or to separate it from silver containing gold, the alloy, which must not contain more than 25 per cent. of gold, is boiled with concentrated sulphuric acid in iron vessels; the concentrated sulphuric acid does not dissolve the iron. The silver and copper are dissolved with the formation of sulphurous anhydride, while the gold remains behind undissolved, as a brown powder. From the solution of silver and copper, the silver is precipitated by copper, and blue vitriol is obtained as a secondary product. This operation is called *refining*.

Sometimes, with the same view, silver containing gold is dissolved in *nitric acid*, which does not dissolve the gold, though it does silver. In this case the remarkable fact is observed, that the silver is completely dissolved only when three-fourths of silver are present to one-fourth of gold (two-thirds of silver, however, is an adequate proportion); hence the term *quartation*. If more than one-fourth or one-third of gold is contained in the alloy, the gold exerts a protecting influence upon the silver, so that the latter is not attacked and dissolved by the nitric acid; then it is necessary to add the proper quantity of silver to the alloy.

The most simple mode of testing gold is to rub some of it off upon a black flint slate (touchstone), and apply to the mark a drop of nitric acid. If the gold is pure, the yellow stroke remains unchanged, but if alloyed, it partly disappears; if it is only an imitation of gold, for instance, tombac, it entirely dissolves.

Gold is represented by the symbol Au, from its Latin name *Aurum*. It forms two classes of salts, typified by aurous chloride, AuCl , and auric chloride, AuCl_3 .

In contact with air, gold is dissolved by solution of potassium cyanide, which converts it to a soluble double cyanide, KAuCy_4 . It is also dissolved by solution of chlorine, usually replaced in practice by *aqua regia*, which acts by virtue of the free chlorine it contains.

Experiment 1.—Put a gold-leaf into a test-tube, and

pour some chlorine solution on it. The gold will soon disappear.

Experiment 2.—Put pieces of gold-leaf into two beakers; pour nitric acid into one, and hydrochloric acid into the other. If the hydrochloric acid contains no free chlorine the gold remains unacted upon in each case. Now mix the contents of the beakers, and a rapid solution of the metal will take place.

Auric chloride, AuCl_3 .—This compound combined with hydrochloric acid, as $\text{AuCl}_3, \text{HCl}$, is formed in both the above experiments, and is the most important of the gold salts. The rest are obtained from it.

Experiment 3.—Heat a few drops of solution of auric chloride on a porcelain-crucible cover, or in a basin. The salt will decompose, and the space it occupied will be covered with a film of metallic gold. All the compounds of gold may be decomposed by mere heating. If the temperature does not exceed 185°C . (365°F .), only part of the chloride is driven off, and *aurous chloride*, AuCl , is formed.

Experiment 4.—Saturate filter-paper with auric chloride solution, dry and burn it completely. Finely divided gold mixed with the ashes of the paper is thus obtained. Rub some on a bright silver spoon with a soft cork, which has been moistened with salt water, and the silver will become gilt (cold gilding). There are other methods of gilding, such as the *moist gilding*, in which the copper, brass, or silver articles are boiled with a diluted solution of gold mixed with sodium hydrogen carbonate; the *hot*, or *mercury gilding*, by which metals are coated with a solution of gold in mercury, and afterwards heated, and the *electro-gilding*, in which the metal is deposited by a galvanic current.

Experiment 5.—Add solution of ferrous sulphate to an acid solution of auric chloride. The mixture immediately assumes a changeable dark and brownish colour, but it appears of a beautiful blue colour on looking through it. On standing, a brown substance is deposited, which is *gold* in the minutest state of division. The ferrous salt at the same time becomes a ferric salt. The finely divided gold is used in the arts for many purposes.

Glass containing it has a ruby colour, and it is used for gilding porcelain, &c., by being mixed with oil of lavender and painted on the surface.

Experiment 6.—To a solution of gold add some stannous chloride which has been partially converted into stannic chloride by long standing. A purple precipitate of unascertained composition will be produced, which is called *purple of Cassius*. This substance produces the most beautiful purple to glass with which it is fused.

Most of the other compounds of gold are unimportant. Gold forms double salts with those of the alkali metals, the most important of which are the double cyanides, from their application in electro-gilding. Two oxides exist; a blue-brown *aurous*, and brown *auric* oxide ($\text{Au}_2\text{O} + \text{Au}_2\text{O}_3$). Hydrogen precipitates a sulphide of gold from an acid solution of auric chloride.

GROUP ii.—Arsenic, Antimony, Bismuth.

ARSENIC.

As = 75. Formula as gas, As_4 .

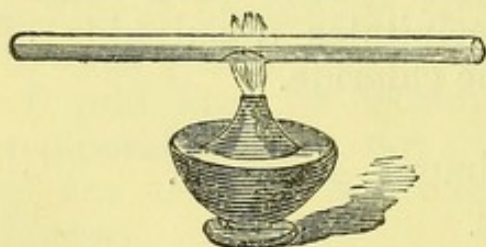
Metallic arsenic is not unfrequently found in the earth, as a lead-grey ore, of strong metallic lustre. It is obtained on a large scale,—*a.*) as a secondary product in the roasting of tin, silver, and cobalt ores; *b.*) as a principal product, by heating arsenical ores with access of air, by which the arsenic becomes oxidised. In both cases the arsenious oxide passes off as vapour, with the smoke, which is conducted through long chambers, and the arsenious oxide condenses as a powder (*white arsenic*, the *arsenic* of the shops). White arsenic is often resublimed in some appropriate apparatus, and is then obtained as *amorphous* arsenious oxide, in solid transparent pieces. These after a time become opaque and *milk-white*, like porcelain, without changing their constitution.

The oxide may then be reduced by charcoal, or the metal is obtained directly by heating *mispickel*, a native compound of iron, with sulphur and arsenic (FeSAs), when the arsenic, being volatile, is driven off in vapour and suitably condensed, $4\text{FeSAs} = 4\text{FeS} + \text{As}_4$.

Metallic arsenic, which possesses a steel-grey lustre at first, soon tarnishes and assumes motley colours in moist air, and finally falls into a coarse grey powder, which is kept on hand in the apothecaries' shops, under the name of fly-poison. If boiled with water, the film of oxidised arsenic dissolves, and a very poisonous liquid is obtained (fly-water). A fresh film of oxide is produced upon the metal which remains, and thus is very easily explained why, after a time, a new poisonous solution can again be prepared from it, without any perceptible decrease of the original powder.

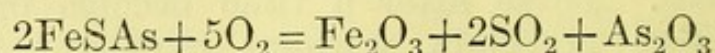
Experiment 1.—Put a small piece of arsenic into a glass

Fig. 89.



tube closed at one end, and heat it; the arsenic volatilises, and deposits itself on the upper portion of the tube as a *brilliant black mirror*; the smell of garlic, peculiar to the fumes of arsenic, being at the same time given off.

Arsenious oxide, or anhydride, As₂O₃.—*Experiment 2.*—Let the arsenical mirror obtained in the above experiment be heated once more, but in an *open* tube; it is converted into a vapour, which condenses on the colder parts of the tube, partly in small white crystals, partly as powder. Before the magnifying-glass the crystals appear as four-sided double pyramids (octahedrons). They consist of *arsenious oxide, or white arsenic*. When arsenic is spoken of in a popular sense this compound is always meant. The preparation of arsenious oxide on the large scale has just been described. The ore most frequently employed is mispickel, all the elements of which are oxidised when the ore is roasted:



Arsenious oxide is soluble in water, but not to a great extent, but it is sufficiently soluble to render these solutions exceedingly dangerous poisons. White arsenic is generally employed for killing rats, moles, and other troublesome house or field animals; for this purpose coloured arsenic only should be purchased, as the white

arsenic looks very much like sugar or flour, and might easily be mistaken for them.

Arsenious oxide prevents the decay of organic substances; therefore the skins of animals intended for shipping are sometimes rubbed with it upon the flesh side.

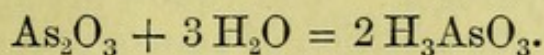
Arsenious oxide when heated readily gives up its oxygen to other bodies; for this reason it is added by glass-makers to melted bottle-glass, to convert its black or green colour into yellow. It acts like black oxide of manganese; namely, it converts the iron from ferrous into ferric salt. On the other hand, it may be made to take up more oxygen and pass into arsenic anhydride.

Experiment 3.—Heat a little arsenious oxide in the bottom of a dry test-tube. It will volatilize and condense as a ring of minute brilliant crystals on the upper part of the tube. It may be noticed that the oxide will volatilize entirely without previously melting.

Experiment 4.—Mix the arsenious oxide with dry powdered charcoal, and then heat in a test-tube. Instead of transparent crystals a brownish mirror of metallic arsenic will be produced, the reduction being effected by the charcoal.

Experiment 5.—If ten grains of arsenious oxide and twenty grains of potassium carbonate are heated with half an ounce of water, the oxide very readily dissolves, and a solution of hydrogen potassium arsenite, KH_2AsO_3 , is obtained.

Arsenious oxide is therefore the anhydride of the tribasic arsenious acid, H_3AsO_3 : it is hence often called arsenious anhydride, and its solution in water is regarded as one of arsenious acid:



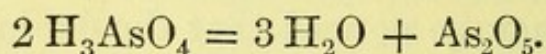
In the last experiment a salt is formed, in which one atom of the hydrogen in this acid is replaced by one of potassium. Its solution is known in medicine as Fowler's solution.

Experiment 6.—Add some of the hydrogen potassium arsenite to a dilute solution of copper sulphate. A bright green precipitate is produced. When dried it forms the

pigment known as *Scheele's green*, often employed for colouring wall-paper. It consists of hydrogen copper arsenite, HCuAsO_3 , a compound in which two of the hydrogen of arsenious acid are replaced by one of the diad copper.

Experiment 7.—Add some silver nitrate solution to another portion of arsenite. A pale yellow precipitate of silver arsenite, Ag_3AsO_3 , is produced; all three of the hydrogen atoms being replaced by silver.

Arsenic oxide, or anhydride, As_2O_5 .—*Experiment 8.*—Boil some arsenious oxide with strong nitric acid in a basin. A violent evolution of red fumes takes place. By evaporation a white deliquescent mass is obtained, which on being heated strongly furnishes arsenic oxide. The deliquescent mass is *arsenic acid*, H_3AsO_4 ; but on being heated it loses its water and becomes the anhydride:



The latter is again converted into arsenic acid by the addition of water. Arsenic acid forms numerous compounds, some of them being found native. It possesses a strong sour taste, and dissolves zinc and iron with liberation of hydrogen, and is in other respects a powerful acid.

Arsenious sulphide, As_2S_3 .—*Experiment 9.*—Dissolve some grains of arsenious oxide in dilute hydrochloric acid, and add to the solution hydrogen sulphide; a precipitate of yellow *arsenious sulphide* is formed, three atoms of sulphur replacing three atoms of oxygen. In this way arsenic may easily be detected in liquids, and separated from them. The salts of cadmium and tin are the only ones, except arsenic, which give a yellow precipitate with sulphuretted hydrogen. The yellow precipitate of arsenic is dissolved by the addition of ammonium sulphide, but is insoluble even in boiling hydrochloric acid.

Arsenious sulphide also occurs native, and is called *orpiment*, or king's yellow, and is used as a yellow pigment.

Experiment 10.—Add hydrogen sulphide to a solution of arsenic acid. No immediate effect is produced, but on heating the mixture, adding more hydrogen sulphide, and allowing it to stand, a yellow precipitate of a lighter colour

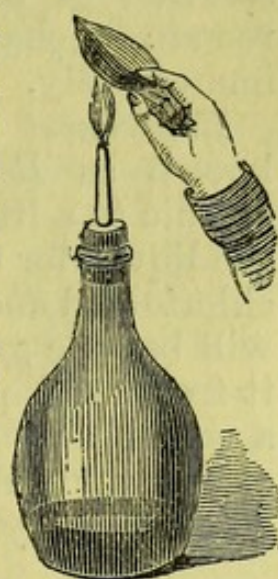
than in the preceding experiment is formed. This is a mixture of arsenious sulphide and free sulphur.

Arsine, or Arseniuretted hydrogen, AsH_3 .—This gas, which is similar in constitution to ammonia and phosphine, cannot easily be obtained pure, but mixed with hydrogen it is evolved in the following experiment:

Experiment 11.—Introduce into a small flask diluted sulphuric acid and some pieces of zinc, and let the hydrogen which is evolved escape through a tube drawn out to a jet, and *after some time* ignite it; you obtain in this manner a hydrogen-lamp. If you hold a glazed porcelain capsule for a few moments in the flame, you will perceive upon it only a circle of small drops of water, which form during the combustion of the hydrogen, and condense on the cold portion. If you now introduce a little arsenious oxide into the flask, the flame, after the gas has been rekindled, will present a bluish-white appearance, and will deposit on the porcelain held in it a *black or brown shining spot* (mirror); this mirror is metallic arsenic. The flame is cooled by a cold body below the temperature which arsenic requires for burning; hence the latter condenses on the porcelain, just in the same way as carbon or soot is deposited on it when held in the flame of a candle. The carbon separates as a light, pulverulent body, arsenic as a coherent mirror. This extremely sensitive test is called, after its inventor, *Marsh's arsenic test*. Care should be taken not to inhale the escaping gas, particularly the unburnt gas; and more than ordinary caution is necessary, as arseniuretted hydrogen is a *most poisonous gas*, and one to which some chemists have already fallen victims.

If a grain or two of arsenious oxide has been introduced into the flask, a white smoke will be seen issuing from the hydrogen flame when it burns freely. This is arsenious oxide, which has been formed by the combination of the arsenic with the oxygen of the air. If a dry test-tube be held with its mouth a little above the flame, so that the

Fig. 90.



smoke may enter, a white sublimate condenses on the interior of the tube. It may be dissolved off with boiling water, and if hydrochloric acid and hydrogen sulphide are then added to the solution, the production of yellow arsenious sulphide will prove the presence of arsenic.

Experiment 12.—Repeat the former experiment, substituting tartar emetic for arsenious oxide; black spots are in this case also deposited on the porcelain, but they are darker, and often have a sooty appearance; they consist of *metallic antimony*. To distinguish spots of antimony with certainty from spots of arsenic, drop upon them a solution of bleaching powder; the spots of antimony remain unchanged, while the arsenical mirrors dissolve immediately.

Experiment 13.—Another excellent test for arsenic is known as *Reinsch's test*. Add a like solution of white arsenic to a little porter or a cup of coffee. Boil some of the liquid for ten minutes with a few drops of pure hydrochloric acid and a small slip of bright copper. The copper will become grey from the deposition of arsenic. Remove it from the liquid, dry it carefully, and heat it gently in a small and *perfectly dry* test-tube. The arsenic will oxidize and sublime in the cool part of the tube in brilliant octohedral crystals easily visible with a lens.

ANTIMONY.

Sb = 120.

Antimony exists native in small quantities, but chiefly in the sulphide, Sb_2S_3 , which constitutes the mineral called *stibnite*, or black antimony ore. This ore is generally called "antimony" in commerce, the metal being known as "regulus of antimony." To obtain the antimony from this ore the ore is melted with pieces of iron, or is roasted to oxide and reduced with carbon. In the former case the iron combines with the sulphur, forming ferrous sulphide, and the antimony is set free.

Antimony has a lamellar crystalline texture, and a white metallic lustre like bismuth, but without the red tint of the latter; it is brittle and may be easily rubbed to powder in a mortar. The sp. gr. of antimony is 6.8. It melts at 425°C . (797°F .)

Of the alloys which antimony forms with other metals, that with lead, with which types are cast, deserves especial notice. Lead alone is much too soft to be employed for this purpose, but if from a fifth to a sixth part of antimony is mixed with it, it acquires such a degree of hardness, that types cast from it may be used for printing many thousand times without losing their sharp edges. A sample of English type metal contains 50 of lead, 25 of antimony, and 25 of copper. From the property possessed by the alloy of expanding as it solidifies, it reproduces the most minute patterns with perfect fidelity. The metal itself, in the free state, is not applied to any important purposes. Its Latin name is *stibium*, whence its symbol Sb.

Antimonious oxide, Sb_2O_3 .—Experiment 1.—Antimony does not alter in the air, but if a piece of it is heated on charcoal before the blow-pipe, it soon melts, and burns with a white flame, forming this oxide, which partly escapes as a white vapour, and is partly deposited as a coating on the charcoal. If you let the melted metallic globule slowly cool, the oxide condenses into crystals, which form around the metal an *espalier* of white points. When thrown into a paper capsule, the white glowing globule will burst into a multitude of small grains, which skip about for some time, leaving in their trail a pulverulent oxide. In the presence of a free supply of air the oxide formed is, however, not Sb_2O_3 , but a higher intermediate oxide, Sb_2O_4 . Antimony generally contains traces of arsenic; hence the smell, like that of garlic, which almost always accompanies its fusion.

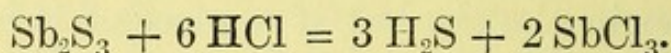
Antimonious oxide may be prepared in a pure state by throwing the corresponding chloride into boiling solution of sodium carbonate, and washing and drying the precipitate.

Antimonic oxide, Sb_2O_5 .—Experiment 2.—Pour a little strong nitric acid containing a little hydrochloric acid on some fragments of antimony. An action similar to that with tin takes place, and the antimony is converted into a white powder. On being dried and gently heated it furnishes antimonic oxide. Combined with water it forms *antimonic acid*, HSbO_3 ; it is therefore properly antimonic

anhydride. The white powder obtained by the nitric acid and antimony is, indeed, antimonious acid, but this is decomposed by the subsequent heating into water and the oxide.

Experiment 3.—Mix some powdered antimony with nitre and throw it into a porcelain crucible which is made red-hot with a Bunsen's burner. *Potassium antimoniate*, KSbO_3 , is thus formed, and may be dissolved out from the fused mass with boiling water.

Antimonious chloride, SbCl_3 .—*Experiment 4.*—Put half an ounce of antimonious sulphide into a capacious flask; pour over it two ounces and a half of strong hydrochloric acid, and heat it in a sand-bath, at first moderately, but afterwards to boiling. Sulphuretted hydrogen escapes, and the chloride is formed:



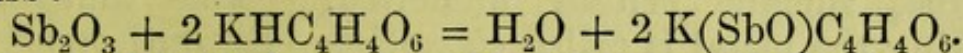
By filtering and evaporating the liquid, antimonious chloride is obtained as a soft yellow solid, which was formerly known as *butter of antimony*. It may be dissolved in a little water and hydrochloric acid, and on then being thrown in a large quantity of water it is converted into a white powder, an oxy-chloride of antimony, SbOCl , similar in constitution to the oxy-chloride of bismuth.

It has already been shown that powdered antimony takes fire if thrown into chlorine gas, and antimonious chloride is formed as a white powder. If more than sufficient chlorine is present to form this compound, *antimonic chloride* SbCl_5 , is produced. It is a yellow fuming liquid, which is decomposed by heat into antimonious chloride and free chlorine.

Potassio-antimonious tartrate, $\text{K}(\text{SbO})\text{C}_4\text{H}_4\text{O}_6$.—This is the best known and most important antimony salt. It is called tartar emetic, and may be regarded as a tartrate of potassium and the monad radical SbO . Oxy-chloride of antimony, SbOCl , is a chloride of the same radical.

Experiment 5.—Boil in a porcelain dish two ounces of distilled water, and during the boiling stir in a mixture of one drachm of antimonious oxide, and one drachm of cream of tartar (hydrogen potassium tartrate, $\text{KHC}_4\text{H}_4\text{O}_6$). When the liquid is half boiled away, filter it while boiling,

and pour one half of it into one ounce of strong alcohol, but set the other half aside. In both cases you obtain *tartar emetic*; in the latter case in the form of crystals, but in the former as a fine powder, because tartar emetic is insoluble in alcohol, and consequently is precipitated by it instantly from its solutions. The following reaction occurs :



We shall learn hereafter that the organic acid tartaric acid, $\text{C}_4\text{H}_6\text{O}_6$, is most simply regarded as containing the diad radical $\text{C}_4\text{H}_4\text{O}_6$ combined with hydrogen, for two of its atoms of hydrogen can be replaced by metals. The name, tartar emetic, indicates the medicinal application of this double salt; it is a very usual means of inducing vomiting. Dissolved in sherry, it forms the well-known *wine of antimony*. It is sparingly soluble in cold water. In large doses it is very poisonous.

Stibine, or *Antimoniuretted hydrogen*, SbH_3 , has already been noticed under arsenic.

Antimonious sulphide, Sb_2S_3 .—*Experiment 6*.—Add some sulphuretted hydrogen to a solution of antimonious chloride, or tartar emetic, in dilute hydrochloric acid; an *orange coloured precipitate* of antimonious sulphide is obtained, which becomes darker on drying. Thus the compounds of antimony may be recognised, as no other metal yields a sulphide of this colour.

The native sulphide, generally known as “black antimony,” or “antimony,” has already been noticed. It is identical in composition with, though so different in appearance from, the sulphide obtained by precipitation.

Many other antimonious and antimonie compounds are known, but few of them are of any interest.

BISMUTH.

Bi = 208.

Bismuth chiefly occurs in the metallic state imbedded in quartz, and in less abundance as sulphide and oxide.

It is also frequently present in cobalt ore, and, in the smelting of this ore for smalt, it separates as cobalt speiss, nickel also being generally present. The metal is

procured from the native ore by a very simple process. It occurs both in the ores and in the speiss in a nearly pure state, and as it melts at a low temperature ($270^{\circ}\text{C.} = 578^{\circ}\text{F.}$), it is only necessary to heat the ores moderately in inclined tubes, when the bismuth melts and flows off below, while the other metals or ores, together with the gangue, remain behind unmelted. This method of working the metal is called *eliquation*. Bismuth is brittle, has a crystalline laminated texture, and a reddish-white colour. Its specific gravity is 9.8.

On allowing melted bismuth to cool till it is covered with a crust, piercing the crust, and then pouring off the fluid portion, the remainder will be found beautifully crystallized in hollow crystals nearly cubical in form.

Experiment 1.—Melt together in a ladle two drachms of bismuth, one drachm of lead, and one drachm of tin. The alloy formed has the very remarkable property of becoming completely liquid when thrown into boiling water. It melts at $94^{\circ}\text{C.} = (201^{\circ}\text{F.})$. By increasing the quantity of lead, alloys may be prepared which readily become liquid at any temperature desired above $212^{\circ}\text{F.} (100^{\circ}\text{C.})$. They are sometimes employed as safety-plates in steam-boilers. The heat of the steam increases with the tension of the steam in the boiler; therefore the alloy to be used has only to be so selected, that, in case of a too great increase of steam, the plate may be melted by the heat of the steam before an explosion of the boiler itself can take place. As this alloy, in its melted state, does not burn the wood, and as it expands during solidification, it is also very well adapted for making metallic copies of engraved wooden moulds for calico-printing and block impressions. This alloy is called *Rose's metal*, after the discoverer.

Bismuth is easily dissolved by dilute nitric acid, but not in hydrochloric acid. It dissolves in heated sulphuric acid.

Bismuthous oxide, or Bismuth tri-oxide Bi_2O_3 .—*Experiment 2.*—Heat a piece of bismuth upon charcoal before the blow-pipe; it melts with the ejection of sparks, and volatilizes at a high temperature with brisk ebullition. A portion of the fumes condense on the charcoal, coating it with a yellow powder; this is bismuth oxide (Bi_2O_3). If you

throw the glowing metallic bead into a small paper box, it divides into small globules, which, while still glowing, will skip about for some moments. An odour like that of garlic, which is frequently emitted during the ignition, proceeds from arsenic, small quantities of which occur in almost all commercial bismuth.

Bismuth tri-nitrate, or Bismuthous nitrate, $\text{Bi}(\text{NO}_3)_3$.—
Experiment 3.—Dissolve some bismuth in dilute nitric acid at a moderate heat, evaporate the solution and set aside to crystallize. The salt will be deposited in large colourless prisms.

Experiment 4.—Redissolve some of the crystals in water to which a little nitric acid is added. Without the addition of acid, solution cannot be effected; but as little as possible must be used. Pour some of the solution into a tumbler of water: an opaque white precipitate of bismuth oxy-nitrate, BiONO_3 , will be produced.

The rest of the bismuth compounds are of little importance.

Bismuthous chloride, BiCl_3 , can be formed by dissolving the oxide in hydrochloric acid and evaporating to dryness, it is a volatile solid and can be easily distilled; on its solution being mixed with water a white precipitate of *bismuth oxy-chloride, BiOCl_2 ,* is produced, similar in constitution to the oxy-nitrate. It is sometimes used as a pigment under the name of *pearl white*.

Another series of bismuth compounds, in which the metal is a pentad, exists. The chief member is bismuth pent-oxide, or bismuthic oxide, Bi_2O_5 ; the rest are little known.

Sulphuretted hydrogen produces a black precipitate of tri-sulphide, Bi_2S_3 , in bismuth solutions. This sulphide occurs native as *bismuthite*.

PART IV.

ORGANIC CHEMISTRY, OR THE CHEMISTRY OF CARBON COMPOUNDS.



CHAPTER I.

INTRODUCTION.

It has already been pointed out that the compounds of carbon are analogous to those of other elements, and that the term "organic," as applied to the former, has no longer the definite meaning it once possessed. It is indeed convenient to retain the old distinction, at least for the present, partly on account of the enormous number and frequent complexity of the carbon compounds, and partly because the constitution of many of them is still unknown to us; but this does not preclude us from adopting in their study the same principles which have guided us through the domain of inorganic chemistry. Carbon itself and a few of its compounds have already been described.

Structure of the simpler carbon compounds.—Let it be remembered that, except in one or two compounds, it is always possible to consider carbon as a tetrad. This is seen plainly in the compounds which have already been noticed, and in a few others of equal simplicity; for example, in carbonic anhydride, CO_2 ; carbon disulphide, CS_2 ; carbon tetrachloride, CCl_4 ; and methane, CH_4 . This is the fundamental principle in the classification of carbon compounds.

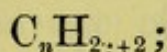
Structure of the Hydrocarbons, or compounds of carbon and hydrogen.—The study of the very numerous compounds of carbon and hydrogen throws a most important light on

the constitution of the more complex carbon compounds and is therefore taken as the starting point in modern organic chemistry. One of these compounds, methane, or marsh-gas (CH_4), contains but one atom, the rest from two to sixteen or more atoms of carbon. The quantity of hydrogen varies from two atoms in acetylene (C_2H_2), up to thirty-four atoms, or even more. Now in examining the formulæ of these hydrocarbons, we are struck with two peculiarities:

1. *The number of hydrogen atoms in a molecule is always an even number.*—We are acquainted, for example, with C_2H_6 , C_2H_4 , and C_2H_2 , but not with C_2H_5 , C_2H_3 , or C_2H .

2. *The number of hydrogen atoms in a molecule is never greater than twice plus two the number of carbon atoms.*—For example, no compound containing two atoms of carbon is known which contains more than six atoms (that is $2 \times 2 + 2$) of hydrogen. And no compound containing seven atoms of carbon is ever found to contain more than sixteen atoms ($7 \times 2 + 2$) of hydrogen.

Saturated Hydrocarbons.—Hydrocarbons which contain the utmost possible quantity of hydrogen, as indicated by the above rule, are called *saturated* hydrocarbons. Their constitution can be expressed by the following very important general formula:



in which n may stand for any number from 1 upwards. Methane, CH_4 , is the type of the class. They are stable compounds, and are, by their nature, incapable of combining directly with chlorine, or any other monad element. This is readily seen in the case of methane, for if CH_4 could combine with Cl and form CH_4Cl , it is evident that carbon in the new compound must be a pentad.

Unsaturated Hydrocarbons.—All hydrocarbons which contain less than the maximum quantity of hydrogen may be described as unsaturated. Ethylene, C_2H_4 , and acetylene, C_2H_2 , for instance, are unsaturated, because one contains two and the other four atoms of hydrogen short of the maximum.

Homologues.—Homologous Series.—The general formula,

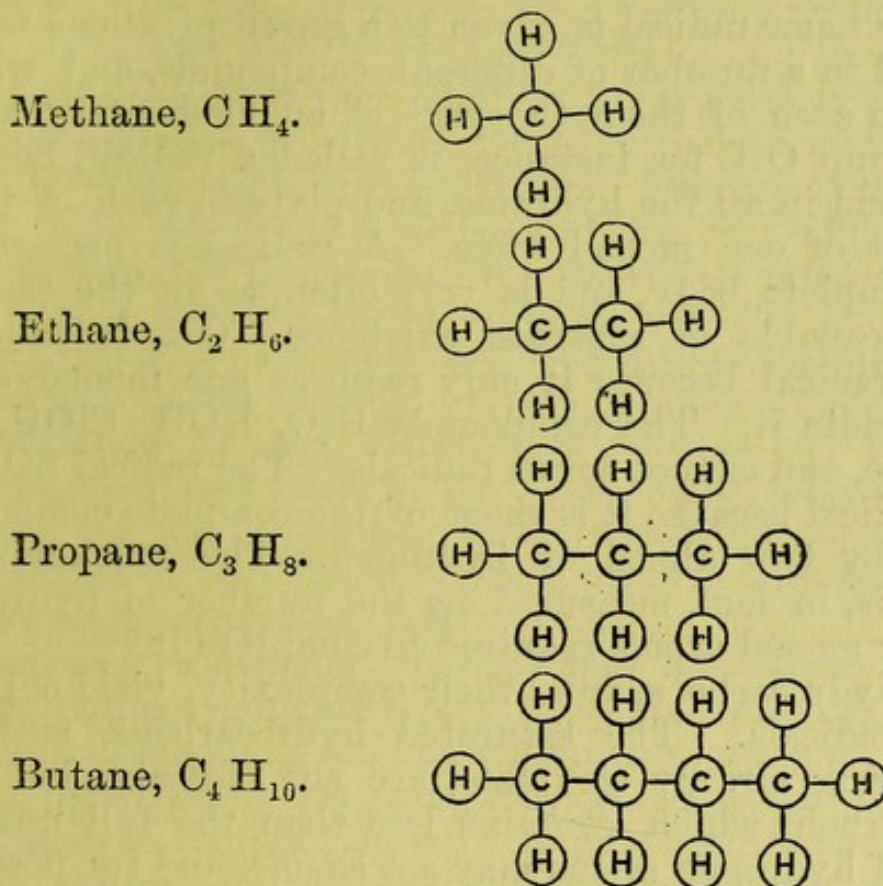
C_nH_{2n+2} , for saturated hydrocarbons will, if studied, bring to light the curious fact that all the members of this series must differ from one another by CH_2 , or by some multiple of CH_2 . The following table, which contains the names and formulæ of a few of the simpler hydrocarbons, shows that this is true not only of this particular series, but also of other series of similar hydrocarbons :

Saturated.	Unsaturated.	
(Methane Series.) C_nH_{2n+2} .	(Ethylene Series.) C_nH_{2n} .	(Acetylene Series.) C_nH_{2n-2} .
CH_4 Methane.	CH_2 Methylene (unknown).	
C_2H_6 Ethane.	C_2H_4 Ethylene.	C_2H_2 Acetylene.
C_3H_8 Propane.	C_3H_6 Propylene.	C_3H_4 Allylene.
C_4H_{10} Butane.	C_4H_8 Butylene.	C_4H_6 Crotonylene.
C_5H_{12} Pentane.	C_5H_{10} Amylene.	C_5H_8 Valerylene.

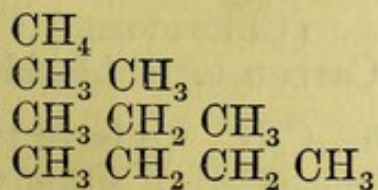
Series of this kind are called *homologous* series, and the members of them are said to be *homologues* of one another. Amylene, for instance, is a homologue of butylene and of ethylene. The phenomenon is not confined to hydrocarbons. Many analogous series are known of compounds which contain oxygen, nitrogen, and other elements in addition to carbon and hydrogen. The arrangement of carbon compounds in series of this kind facilitates their study to an amazing extent, for the members of each series generally resemble one another so closely as to make it unnecessary to study more than one or two of them.

Hypothetical explanation of the foregoing fact.—By a slight extension, the doctrine of valency affords an intelligible explanation of these peculiarities of carbon compounds. Let us revert to the illustration employed before, and suppose that every elementary atom has one or more *bonds* or *arms* by which it can grasp and be grasped by other atoms. The monads have one bond, the diads two, the triads three, and the tetrads four. We can then conceive how the carbon atom with its four bonds can attach to itself four atoms of hydrogen. And we can understand further that when there are two atoms of carbon, these two must be linked together by at least one bond of each, so that there will remain only six free bonds by which hydrogen atoms can be attached. In like manner, three

carbon atoms can only combine with eight hydrogen atoms, because, for three atoms to be linked together one of them must use two of its bonds. The following diagram will show how simply this conjecture explains the constitution of the saturated hydrocarbons :



The unsaturated hydrocarbons, and indeed all chemical compounds the nature of which is understood, can be represented by formulæ analogous to these "glyptic" formulæ. But their use is quite unnecessary, for it is sufficient to write against each carbon atom the symbols for the elements or radicals which are attached to it. Thus the above compounds are generally represented by the following, which are called *structural formulæ* :



It is well to repeat that no one supposes that the

atoms really have anything like arms projecting from them. The way in which atoms are held together is unknown to us, and the hypothesis only pretends to supply one conceivable mode of accounting for the structure of compounds.

Hydrocarbon Radicals.—It has already been explained that the name radical is given to a group of atoms which is found in a number of different compounds, and which plays in each of them the part of an elementary atom. The group OH, for instance, is called a radical, because it is found in all the hydrates, and plays in each of them the part of one monad atom. A radical is necessarily an incomplete body, and is very often, as in the case of OH, incapable of separate existence. OH is called a monad radical because it only requires one monad atom to complete it. The compounds H_2O , KOH , ClOH , are complete, and cannot act as radicals. The radical SO_4 is a diad radical because it is short of the complete compound H_2SO_4 by two atoms of hydrogen. The valency of a radical is, in fact, measured by the number of hydrogen or other monad atoms required to complete it.

The hydrocarbons, from their complexity, yield a great many radicals. The saturated hydrocarbons, such as CH_4 , being already complete, are not radicals; but any hydrocarbon which contains less than the full number $2n+2$ of hydrogen atoms may act as one, and its possible valency may be measured by the number of hydrogen atoms which it lacks. Thus we have the following radicals derived from methane. The chlorine compound is written after each to illustrate the kind of compound into which it enters.

	CH_4	Methane.		
Radicals.	$(\text{CH}_3)'$	Methyl.	CH_3Cl	Methyl chloride.
	$(\text{CH}_2)''$	Methylene.	CH_2Cl_2	Methylene dichloride.
	$(\text{CH})'''$	Methenyl.	CHCl_3	Methenyl trichloride. (Chloroform.)
	C^{iv}	Carbon.	CCl_4	Carbon tetrachloride.

In the same way, related to ethane, C_2H_6 , we have the radicals $(\text{C}_2\text{H}_5)'$ ethyl, $(\text{C}_2\text{H}_4)''$ ethylene, $(\text{C}_2\text{H}_3)'''$ ethenyl, $(\text{C}_2\text{H}_2)^{\text{iv}}$ acetylene, and $(\text{C}_2\text{H})^{\text{v}}$. The greater the number

of carbon atoms, the greater of course, will be the possible number of hydrocarbon radicals in the series.

Any hydrocarbon radical which contains an even number of hydrogen atoms is capable of separate existence but the others are not, and are therefore to be compared with OH, NH_4 , &c. In the series last named, C_2H_6 , C_2H_4 , and C_2H_2 are well known.

Valency of Hydrocarbon Radicals.—The rule just given, namely, that the valency of the radical may be inferred from the number of hydrogen atoms which it requires to bring it to the formula $\text{C}_n\text{H}_{2n+2}$, is not universally true. The radical phenyl, C_6H_5 , for instance, should by this rule have a valency equal to nine, whereas it is only found to be a monad; and oil of turpentine, $\text{C}_{10}\text{H}_{16}$, which should be a hexad, is only a tetrad. Many hydrocarbons which should by their formulæ be radicals do not fulfil this function, and seem to be complete, or *saturated*. It is easy by the hypothesis before described to give an explanation of these peculiarities, but it is only necessary in this place to note the fact.

Compounds containing Hydrocarbon Radicals.—Hydrocarbon radicals generally play in compounds the part of metals. Compounds are known which correspond closely to the hydrates, oxides, salts, &c., of the metals, differing from them in composition only by containing some hydrocarbon radical in place of the metal. Many of these compounds are among the most important in chemistry. It is essentially necessary to understand and remember their formulæ, and this will be done with ease if their resemblance to the metallic compounds is carefully borne in mind.

To begin with the monad radicals, such as methyl, CH_3 , ethyl, C_2H_5 , &c. They are strictly comparable to atoms of potassium. Take the formula of any potassium compound, replace the symbol K by the group CH_3 , and you will have the formula for a *methyl* compound—oxide, chloride, hydrate, or oxy-salt—which shall resemble the potassium compound in many particulars.

So again with the compounds of diad radicals. The diad radical ethylene, C_2H_4 , is like an atom of zinc, and the ethylene compounds are curiously similar to those

of zinc. The following table will supply a few illustrations of this interesting resemblance :

CH_3Cl	Methyl chloride	like	KCl	Potassium chloride.
CH_3OH	hydrate	"	KOH	hydrate.
	(Methyl-alcohol.)			
$(\text{CH}_3)_2\text{O}$	oxide	"	K_2O	oxide.
	(Methyl-ether.)			
CH_3NO_3	nitrate	"	KNO_3	nitrate.
CH_3HSO_4	{hydrogen sulphate.}	"	KHSO_4	{hydrogen sulphate.
$(\text{CH}_3)_2\text{SO}_4$	sulphate	"	K_2SO_4	sulphate.
$\text{C}_2\text{H}_4\text{Cl}_2$	Ethylene chloride	"	ZnCl_2	Zinc chloride.
$\text{C}_2\text{H}_4(\text{OH})_2$	hydrate	"	$\text{Zn}(\text{OH})_2$	hydrate.
	(Glycol.)			
$\text{C}_2\text{H}_4\text{O}$	oxide	"	ZnO	oxide.
$\text{C}_3\text{H}_5\text{Cl}_3$	Propenyl chloride,	"	BiCl_3	Bismuth chloride.
$\text{C}_3\text{H}_5(\text{OH})_3$	hydrate	"	$\text{Bi}(\text{OH})_3$	hydrate.
	(Glycerine)			
$\text{C}_3\text{H}_8\text{Br}_4$	Valerylene bromide	"	SnBr_4	Stannic bromide.

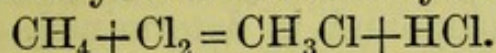
Nomenclature of the above compounds.—Besides the systematic names, founded on their analogy with metallic compounds, many of these compounds are known by other names, which were given to them before their nature was understood. Thus the hydrates are generally called *alcohols*, common alcohol, $\text{C}_2\text{H}_5\text{OH}$, being a member of the class. Alcohols are sometimes said to be *monatomic*, *diatomic*, or *triatomic*, according as they are the hydrates of monad, diad, or triad radicals. Diatomic alcohols are also known as *glycols*, from the old name of ethylene hydrate, $\text{C}_2\text{H}_4(\text{OH})_2$, and triatomic alcohols, as *glycerols* or *glycerines*; common glycerine, $\text{C}_3\text{H}_5(\text{OH})_3$, being the only important member of the class.

The oxides of the radicals are called *ethers*, from common ether, which is the oxide of ethyl, $(\text{C}_2\text{H}_5)_2\text{O}$. The salts—chlorides, sulphates, &c.—are sometimes, but very improperly, called *compound ethers*.

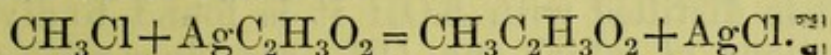
Modes in which the above compounds are formed.—These compounds are produced by a great variety of processes. Some of them are formed by the processes of animal or vegetal life; as, for instance, the natural fats and oils, which are salts of the triad radical propenyl: some by artificial processes, the nature of which is as yet imperfectly understood, such as fermentation and destructive

distillation. Very many can be prepared artificially, either from hydrocarbons or from other carbon compounds, and in these cases a satisfactory account can generally be given of their mode of formation. The processes in fact often bear a close resemblance to those of mineral chemistry. By way of illustration, let us take a few of the derivatives which may be obtained from marsh-gas, or methane.

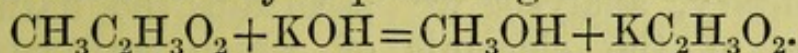
(1.) Methane mixed with chlorine and exposed to sunlight yields methyl chloride and hydrochloric acid :



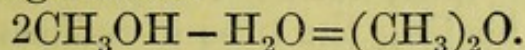
(2.) Methyl chloride with silver acetate yields methyl acetate and silver chloride :



(3.) Methyl acetate with potassium hydrate yields methyl hydrate (methyl alcohol) and potassium acetate. This is a common way of producing alcohols :



(4.) Methyl hydrate, dehydrated by sulphuric acid, yields methyl oxide (methyl ether). This is a common way of producing ethers :



Isomerism, Metamerism, Polymerism.—It often happens that two or more compounds contain the same elements in the same proportions, and even contain in each molecule the same number of atoms, the difference between them consisting entirely in the arrangement of the atoms. If in such compounds all the carbon atoms are united directly together, the compounds are said to be *isomeric*. Thus CH_2Cl , CH_2Cl_2 and CH_3CHCl_2 are isomeric though not identical. But if two or more carbon radicals are linked together by a dissimilar atom, such as O, S, or N, the compounds are said to be *metameric*. Thus $\text{CH}_3\text{CH}_2\text{OH}$ (ethyl alcohol) and CH_3OCH_3 (methyl ether) are metameric, because in the second compound the two carbon atoms are separated and yet united by an oxygen atom.

Polymeric bodies are those in which the molecular weights are multiples of one another, as, for example, ethylene, C_2H_4 , and butylene, C_4H_8 .

Analysis of Carbon compounds.—The analysis of organic substances is of two kinds—proximate and ultimate. Proximate analysis is applicable in the first place to mixtures. It teaches us what compounds are mixed together in any given substance. Coal tar, for instance, is a highly complex mixture. Subjecting this to a variety of processes of proximate analysis we find out that it contains water, various basic and acid bodies, neutral hydrocarbons, &c. The same term is sometimes applied to many of the processes by which the arrangement of the atoms in a compound is determined. The substance called methyl acetate, for example, is found, by ultimate or elementary analysis, to have the formula $C_3H_6O_2$. But when it is distilled with KOH it is found to divide into two portions, of which one, CH_3 , combines with the OH, and the other, $C_2H_3O_2$, with the K. We therefore ascribe to the compound the formula $CH_3.C_2H_3O_2$. By ultimate, or elementary, analysis we ascertain the nature and proportion of the elements of which any compound consists. Almost all carbon compounds are analysed by mixing a known weight of them with cupric oxide in a hard glass tube, called a combustion tube, closed at one end, and heating the tube to redness in a charcoal or gas furnace. The carbon, taking oxygen from the cupric oxide, becomes CO_2 , and the hydrogen H_2O . The water is absorbed in a tube filled with calcium chloride, the carbonic anhydride in a series of bulbs partly filled with solution of caustic potash. The calcium chloride tube and the potash bulbs are weighed before and after the experiment; the increase of weight showing the quantity of water and carbonic anhydride formed by the combustion of the substance. From these quantities, that of the carbon and hydrogen can easily be calculated. If nitrogen is present it is given off as gas during the combustion, and can be collected and measured. Other elements, such as chlorine, sulphur, metals, &c., can be estimated by appropriate methods. After all other elements have been estimated and calculated as parts per cent., the quantity required to make up 100 must be oxygen, which is always determined in this way.

CHAPTER II.

CLASSIFICATION OF THE SIMPLER CARBON COMPOUNDS.

BEFORE commencing the experimental study of organic chemistry, the student will find it well to master the outlines of the system by which organic compounds are classed. The following are a few of the most important classes :

1. *Hydrocarbons*; many of which are radicals.
2. *Alcohols*, or hydrates of hydrocarbon radicals. These radicals are sometimes called *alcohol radicals*.
3. *Ethers*, or oxides of hydrocarbon radicals.
4. *Salts of hydrocarbon radicals*; sometimes called *compound ethers*.
5. *Organic acids*, or hydrates of acid radicals. Acid radicals are, for the most part, hydrocarbon radicals, in which a portion, or the whole of the hydrogen is replaced by oxygen.
6. *Ammonia derivatives*, in which a portion, or the whole of the hydrogen of ammonia is replaced by alcohol radicals or acid radicals. The *alkaloids* or *organic bases* belong to this class.
7. *Substitution compounds*, in which hydrogen is partly, or wholly replaced by Cl, Br, I, NO₂, &c.

All the above compounds can be arranged in homologous series, each of which can be represented by a general formula, in which n may stand for any number.

Many compounds are still too imperfectly known to be referred with certainty to any of the classes.

HYDROCARBONS.

A few of the simplest compounds of carbon and hydrogen have already been described, and the others will be more conveniently considered in subsequent chapters. A short classification of the most important of them is all that is necessary here.

Series C_nH_{2n+2} ; sometimes called the Paraffins.—Methane, or marsh-gas, CH_4 is the lowest member of this, which is *par excellence* the saturated series. The solid substance called *paraffin* is a mixture of several of the higher members of the series. The petroleum oils which issue from the earth in America and other places, and the so-called paraffin oils, obtained by destructive distillation from some kind of coal, are very complex mixtures of hydrocarbons, belonging to this series. Many of them also can be formed in the laboratory by artificial processes.

Series C_nH_{2n} ; sometimes called the Olefines.—Ethylene, C_2H_4 , described before, is the lowest known member. Many of them can be obtained by dehydrating some alcohol by means of sulphuric acid or zinc chloride. They are also frequently formed during destructive distillation, a process to which a subsequent chapter is devoted. The true olefines are all diad radicals, inasmuch as each of them will combine with two atoms of chlorine.

Series C_nH_{2n-2} .—Acetylene, C_2H_2 (p. 185), is the only important member.

Series C_nH_{2n-4} .—Oil of turpentine, and many other of the volatile oils described in a later chapter, have the formula $C_{10}H_{16}$, or the polymeric formulæ $C_{15}H_{24}$, or $C_{20}H_{32}$, and therefore belong to this series. As the saturated hydrocarbon would have the formula $C_{10}H_{22}$, it is evident that these oils might act as hexad radicals. In practice they only unite with four monad atoms, and are therefore reckoned as tetrads.

Series C_nH_{2n-6} .—Benzene, C_6H_6 , which is found in the most volatile portion of coal tar, is the most important member, but several higher homologues are known.

Series C_nH_{2n-12} .—Naphthalene, $C_{10}H_8$, also found in coal tar, is the only example of any importance.

Series C_nH_{2n-18} .—Anthracene, $C_{14}H_{10}$, is found in the least volatile portion of coal tar.

ALCOHOLS.

Methyl series, $C_nH_{2n+1}OH$.—This, which contains common alcohol, is by far the most important. The following are members of it:

Methyl alcohol, CH_3OH . Formed in the destructive distillation of wood.

Ethyl, or common alcohol, $\text{C}_2\text{H}_5\text{OH}$.

Trityl, or propyl ,, $\text{C}_3\text{H}_7\text{OH}$.

Tetryl, or butyl ,, $\text{C}_4\text{H}_9\text{OH}$.

Pentyl, or amyl ,, $\text{C}_5\text{H}_{11}\text{OH}$.

Hexyl ,, $\text{C}_6\text{H}_{13}\text{OH}$.

Heptyl ,, $\text{C}_7\text{H}_{15}\text{OH}$.

The last six are all formed during the fermentation of sugar, but common alcohol in greatest quantity.

Cetyl alcohol, $\text{C}_{16}\text{H}_{33}\text{OH}$, from spermaceti.

Myricyl alcohol, $\text{C}_{30}\text{H}_{61}\text{OH}$, from bee's-wax.

Allyl series, $\text{C}_n\text{H}_{2n-1}\text{OH}$.—Allyl alcohol, $\text{C}_3\text{H}_5\text{OH}$, an unsaturated alcohol.

Benzyl series, $\text{C}_n\text{H}_{2n-7}\text{OH}$.—Benzyl alcohol, $\text{C}_7\text{H}_7\text{OH}$, which can be prepared from oil of bitter almonds, is the lowest and most important of this series.

Phenols.—A few curious substances are known, of which ordinary phenol (called also phenyl alcohol, and more frequently carbolic or phenic acid) is the type. They exhibit, to a certain extent, the nature both of alcohols and of acids, and are yet different from either. Common phenol, which, under the name of carbolic acid, is largely used as an antiseptic, is formed during many processes of destructive distillation. Its formula is $\text{C}_6\text{H}_5\text{OH}$, and it is a hydrate of the monad radical *phenyl*, C_6H_5 .

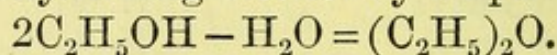
Diatomic Alcohols, or Glycols.—Very few are known, and those have no practical importance. Common glycol is $\text{C}_2\text{H}_4(\text{OH})_2$; it is therefore the hydrate of the diad radical ethylene, C_2H_4 .

Triatomic Alcohols, Glycerols, or Glycerines.—Common glycerine, obtained from most fixed oils and fats, is the only important member of the series. It has the formula $\text{C}_3\text{H}_5(\text{OH})_3$, and is therefore the hydrate of the triad radical propenyl or glyceryl, C_3H_5 .

ETHERS.

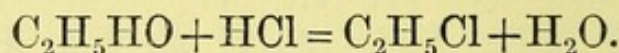
Common ether may be taken as an example of compounds of this class. It is an oxide of the monad radical

ethyl C_2H_5 , the radical of common alcohol, and its formula is $(C_2H_5)_2O$ or $C_4H_{10}O$. The relations of the alcohols and ethers to one another and to metallic hydrates and oxides have already been pointed out. Ethyl oxide, or ether, is prepared by dehydrating alcohol by sulphuric acid:



SALTS.

The salts of hydrocarbon radicals can be prepared by various processes, but the simplest and one of the commonest consists in treating the corresponding hydrate, or alcohol, with the acid. Ethyl chloride, for example, may be prepared by treating alcohol with hydrochloric acid gas:

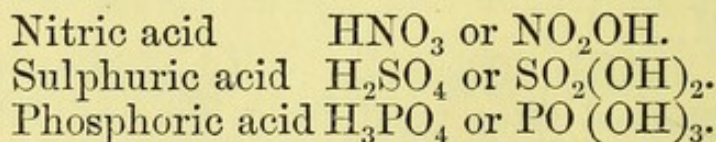


In other cases the alcohol is treated with the materials from which the acid is made. Thus common alcohol with sodium acetate and sulphuric acid gives ethyl acetate (acetic ether) and sodium sulphate.

These compounds are sometimes called "compound ethers."

ACIDS.

Oxygen-containing acids have hitherto been regarded as compounds of hydrogen with radicals analogous in function to chlorine. But it is equally easy and often more convenient to regard them as *hydrates*, that is, as containing the radical OH instead of H. The acid radical will then of course contain less oxygen. The difference between the two views will be readily understood by the following formula for three of the commonest mineral acids:



The student will find no difficulty in applying these formulæ to all known oxy-acids. The radicals of this kind are distinguished by the termination *yl*; thus NO_2 is called *nitryl*, SO_2 *sulphuryl*, and PO *phosphoryl*. It will be seen at a glance that they are respectively monad, diad, and triad.

This view renders the structure of the simpler organic acids easy to understand. Every organic acid may be regarded as the hydrate of an acid radical analogous to the above-named mineral radicals. And this radical is, in all the simpler acids, a hydrocarbon in which part or all of the hydrogen is replaced by oxygen, one atom of oxygen replacing two atoms of hydrogen. Acids and alcohols are therefore related very simply to one another, as the following examples will show :

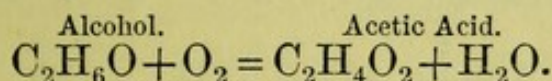
Methyl alcohol, CH_3OH .

Formic acid, CHOOH .

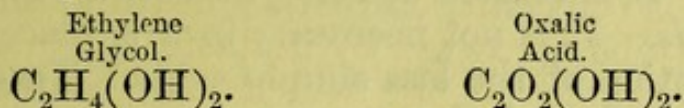
Ethyl alcohol, $\text{C}_2\text{H}_5\text{OH}$.

Acetic acid, $\text{C}_2\text{H}_3\text{OOH}$.

In these and in many other cases the acids can actually be formed from the alcohols by direct oxidation. Thus when dilute alcohol absorbs oxygen from the air it is converted to vinegar, that is, to acetic acid :



Similar relations can be observed between diatomic alcohols and certain dibasic acids, as is shown by the following comparison :



It must, however, be remembered that many acids are known to which no corresponding alcohol has been discovered.

Let it be distinctly understood that either way of writing the formula of an acid may with correctness be employed. It does not much matter for ordinary purposes whether we call acetic acid $\text{HC}_2\text{H}_3\text{O}_2$ or $\text{C}_2\text{H}_3\text{OOH}$.

The following are a few of the most important classes of organic acids. For the sake of comparison with the alcohols, the acids are here formulated as hydrates.

Formic Series—Fatty Acids, $\text{C}_n\text{H}_{2n-1}\text{OOH}$.

This important series is related in the manner above described to the first series of alcohols. The acids belonging to it are often called the fatty acids, because several of the higher members (palmitic, stearic) are

derived from fats and oils, which are indeed glyceryl salts of those acids. The most important acids are given in the following table :

Fatty Acid.		Corresponding Alcohol.
Formic	CHOOH.	Methyl $\text{CH}_3\text{OH.}$
Acetic	$\text{C}_2\text{H}_3\text{OOH.}$	Ethyl $\text{C}_2\text{H}_5\text{OH.}$
Propionic	$\text{C}_3\text{H}_5\text{OOH.}$	Trityl $\text{C}_3\text{H}_7\text{OH.}$
Butyric	$\text{C}_4\text{H}_7\text{OOH.}$	Tetryl $\text{C}_4\text{H}_9\text{OH.}$
Valeric	$\text{C}_5\text{H}_9\text{OOH.}$	Amyl $\text{C}_5\text{H}_{11}\text{OH.}$
Caproic	$\text{C}_6\text{H}_{11}\text{OOH.}$	Hexyl $\text{C}_6\text{H}_{13}\text{OH.}$
Palmitic	$\text{C}_{16}\text{H}_{31}\text{OOH.}$	Cetyl $\text{C}_{16}\text{H}_{33}\text{OH.}$
Stearic	$\text{C}_{18}\text{H}_{35}\text{OOH.}$	

Acrylic Series, $\text{C}_n\text{H}_{2n-3}\text{OOH.}$

Acrylic acid, $\text{C}_3\text{H}_3\text{OOH}$, prepared from glycerin.

Oleic acid, $\text{C}_{18}\text{H}_{33}\text{OOH}$, „ „ olive and other oils.

Benzoic Series, $\text{C}_n\text{H}_{2n-9}\text{OOH.}$

Benzoic acid, $\text{C}_7\text{H}_5\text{OOH}$, prepared from gum benzoin and in other ways, is the lowest and most important of the series. It is related to benzyl alcohol, $\text{C}_7\text{H}_7\text{OH}$.

Other acids.—It is not necessary in this place to describe the constitution of the less simple acids. The names and formulæ of a few of the more important are appended; the formulæ being written on the older and, in some respects, simpler plan :

Monobasic.

Lactic acid, $\text{HC}_3\text{H}_5\text{O}_3$. Found in sour milk and in the juice of flesh.

Gallic acid, $\text{HC}_7\text{H}_5\text{O}_5$. In gall nuts.

Dibasic

Oxalic acid, $\text{H}_2\text{C}_2\text{O}_4$. In wood sorrel and other plants, also by the oxidation of starch, sugar, &c., by nitric acid.

Succinic acid, $\text{H}_2\text{C}_4\text{H}_4\text{O}_4$. By the distillation of amber (succinum), and in other ways.

Malic acid, $\text{H}_2\text{C}_4\text{H}_4\text{O}_5$. In apples and other fruits.

Tartaric acid, $\text{H}_2\text{C}_4\text{H}_4\text{O}_6$. In grapes and other fruits.

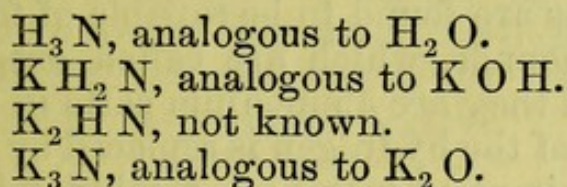
Tribasic.

Citric acid, $\text{H}_3\text{C}_6\text{H}_5\text{O}_7$. In lemons and other fruits.

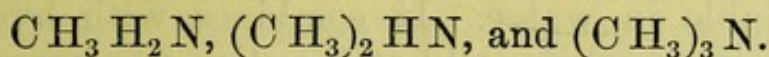
AMMONIA DERIVATIVES.

We have seen that a portion, OH, of water is found in a large class of compounds, and is therefore called a radical. And we have also seen that three several portions of methane, CH_4 , may act as radicals, namely, CH_3 , CH_2 , and CH . In like manner, two portions of ammonia gas, NH_3 , namely, NH_2 and NH , both probably incapable of separate existence, are found to play, in certain compounds, the part of radicals. NH_2 , sometimes called *amidogen*, is of course a monad, like OH and CH_3 ; and NH , sometimes called *imidogen*, a diad. They are formed by replacing one or two atoms of the hydrogen of ammonia by metals, by alcohol radicals, or by the acid radicals of which the acids are hydrates. And it is possible, in many cases, to replace the third atom of hydrogen by a metal or radical. The metal or radical is then combined with nitrogen only.

Thus, starting from ammonia, we have—



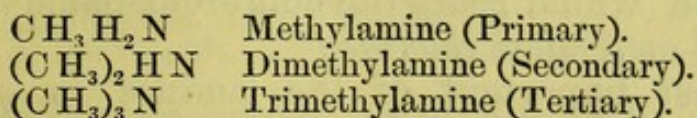
In the same way, with the radical methyl, we have—



The metallic derivatives of ammonia are, for the most part, unimportant, but those which contain alcohol or acid radicals are very interesting. In the nomenclature which is most frequently employed the former are called *amines*, whether they contain NH_2 , NH , or N , and the latter *amides*, and both are known as *primary*, *secondary*, or *tertiary*, according as one, two or three atoms of hydrogen are replaced.

Amines.

1. Of monad radicals.



Which may serve as examples. Amines containing other alcohol radicals are formed and named in the same way.

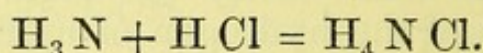
Two or three different alcohol radicals may be present in an amine, as, for example :

$\text{C H}_3 \text{ C}_2 \text{ H}_5 \text{ C}_5 \text{ H}_{11} \text{ N}$, Methyl-ethyl-amylamine.

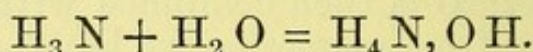
2. Of diad radicals. Generally called diamines.

$\text{C}_2 \text{ H}_4 (\text{H}_2 \text{ N})_2$, called ethylene diamine, is a sufficient example.

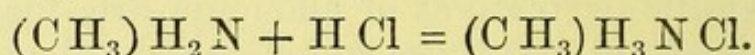
Compound ammoniums.—Amines are similar to ammonia in constitution and generally in properties. Now we know that ammonia combines with hydrochloric and similar acids, and forms compounds which are called ammonium salts :



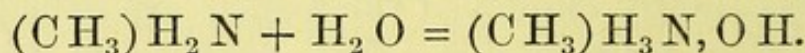
And it is usual to assume that the solution of ammonia in water contains a definite hydrate of the same radical, ammonium :



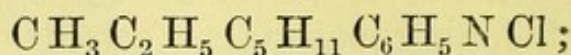
The amines are found to be capable of forming precisely similar compounds, which are called *compound ammonium salts*, because they are ammonium salts in which a portion or the whole of the hydrogen is replaced by alcohol radicals. To take the simplest instance—methylamine, $\text{CH}_3 \text{ H}_2 \text{ N}$, is a gas, very similar to ammonia in properties, and even more soluble in water. It combines with hydrochloric acid, forming a white solid salt, *methyl-ammonium chloride* :



Its solution in water contains the hydrate :



Ammonium salts have been prepared in which all four of the atoms of hydrogen are replaced by the same or by different radicals. For example,



a compound which bears the cumbrous name *methyl-ethyl-amyl-phenyl-ammonium chloride* !

The most important of the amines is *phenylamine*, generally called *aniline*. It contains the radical phenyl, and has the formula $\text{C}_6 \text{ H}_5 \text{ H}_2 \text{ N}$.

Amides.

1. Of monad radicals.

$C_2H_3OH_2N$, Acetamide, is a good example. It contains acetyl, C_2H_3O , the hydrate of which is acetic acid. The analogy between the amide and acid is seen at a glance.

$C_2H_3ONH_2$, Acetyl amide, or acetamide.

C_2H_3OOH , Acetyl hydrate, or acetic acid.

2. Of diad radicals.

$CO(NH_2)_2$, Carbonyl diamide, commonly called *urea*, is an extremely important example. It contains the radical CO of carbonic acid.

$CO(OH)_2$ Carbonic acid ($=H_2CO_3$).

$CO(NH_2)_2$, Urea.

Amides are commonly neutral substances.

Natural bases, or alkaloids.—It cannot be doubted that these very important compounds are derivatives of ammonia, but their exact structure is not yet ascertained. A chapter will be devoted to their study.

SUBSTITUTION COMPOUNDS CONTAINING Cl , NO_2 &c.

In many compounds a portion of the hydrogen may be replaced by chlorine, by the radical NO_2 , or similar bodies, without the character of the compound being entirely lost.

Chloro-, Bromo-, and Iodo- compounds.—We have already seen that chlorine may displace the hydrogen in hydrocarbons. The following compounds, obtained from acetic acid, will explain themselves. Many analogous bodies are known.

$C_2H_4O_2$, Acetic acid.

$C_2H_3ClO_2$, Chloracetic acid.

$C_2H_2Cl_2O_2$, Dichloracetic acid.

$C_2HCl_3O_2$, Trichloracetic acid.

All of these are monobasic acids, like acetic.

Nitro-compounds.—Many substances, when treated with nitric acid, exchange one or more atoms of hydrogen for

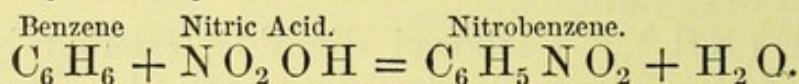
the radical NO_2 . Compounds so formed are called *nitro-compounds*. Examples :

C_6H_6 , Benzene.

$\text{C}_6\text{H}_5\text{NO}_2$, Nitrobenzene.

$\text{C}_6\text{H}_4(\text{NO}_2)_2$, Dinitrobenzene.

Many important explosive agents, for example gun-cotton and nitro-glycerine, are either nitro-compounds or nitrates isomeric with them. The mode in which they are formed is seen very clearly in the case of nitrobenzene :



The modern system of classification has not been employed in the following pages. Great as are the advantages it presents, it does not appear so well fitted for the purposes of a simple experimental study as the older and more arbitrary arrangement here adopted. The student will find it wise, while investigating the properties of each substance he deals with, to mark its formula and endeavour to ascertain its true position among carbon compounds. This will not always be possible, but in very many cases a reference to this chapter will enable him to succeed.

CHAPTER III.

CELLULIN, STARCH, GUM, AND SUGAR GROUP.

THE members of this highly important group of compounds have a complex structure which is as yet but imperfectly understood. They have not been prepared artificially, but are formed by the subtle processes which are connected with the functions of vegetable and animal life. Some of them are peculiar to vegetables, a few to animals, and all are connected together by a great similarity of composition and properties. Some of them, as, for instance, starch and sugar, are highly important constituents of food.

It will be seen, in the following table, that they all consist of carbon, hydrogen, and oxygen, and that the two latter elements are in almost every case present in the same proportions as in water. They are therefore called *carbo-hydrates*. Cellulin and starch are distinguished from the rest, in that they have a definite *structure*—fibres and granules. They are therefore said to be *organized* compounds.

Cellulin, or vegetable fibre.	}	All of which have the formula $C_6H_{10}O_5$.
Starch, in the seeds and other parts of plants.		
Glycogen, in the tissues of animals and in certain vegetables.		
Dextrin, formed by the alteration of starch, &c.		
Arabin, in gum arabic.	}	$C_6H_{12}O_6$.
Glucose, or grape-sugar, found in honey, fruits, &c., also in the animal body. Can also be formed from cellulin, starch, dextrin, &c.		

Sucrose, or cane-sugar, found in the sugar-cane, maple, beet-root, and other plants.	} $C_{12}H_{22}O_{11}$.
Lactose, or milk-sugar, in the milk of animals.	
Maltose, in malt and wort.	

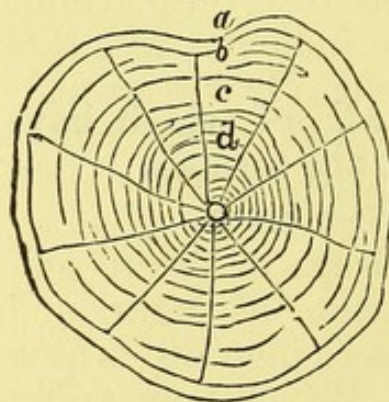
The true formulæ of many or all of these compounds are probably multiples of those here given.

CELLULIN OR CELLULOSE.

All the cells and vessels of plants are composed of cellulin. This substance is to plants what bones, flesh, and skin are to the animal body; it forms the solid mass of all vegetable organs, and consequently imparts to plants their shape and firmness; it forms the ducts or veins of the plants, through which the sap circulates. We find it very finely ramified, tender, soft, and easily digestible in the young leaves, flowers, and stems, and in the so-called pulp of fruits and roots, as apples, plums, carrots, &c.; hard and indigestible in straw, wood (*woody tissue*), and in the husks of grain (*bran*); hardened like stone in the stones of plums, cherries, and peaches, and in the shells of nuts; light, porous, and elastic in the pith of the elder, and in the cork; lengthened and pliant in hemp, flax, and cotton.

The transverse section of the stem of a tree illustrates the influence which age exerts upon the vegetable tissue

Fig. 91.

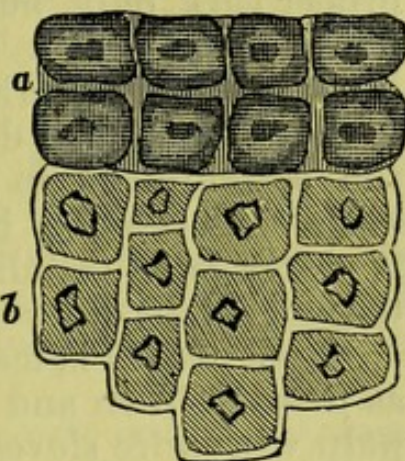


and how this tissue varies in one and the same tree. Inside the *bark* (*a*) lies the *inner fibrous bark* (*b*), which consists of lengthened tubes, and is peculiarly adapted to supply the place of veins in the tree. Here the sap principally circulates, and therefore a tree will die when the inner bark is girdled, whilst (as seen in many hollow trees) the tree will live on if only the inner and outer

bark remain, though the wood itself is entirely rotten and gone. From the inner bark towards the exterior is deposited

every year a new layer of bark, and towards the centre a new layer of wood (annual circles). The light and whiter wood, lying next the inner bark, is called *sap wood* (*c*); but this by the annually increasing compressure, becomes denser and more solid, and then it is called *heart-wood* (*d*). The latter is usually darker, and is frequently impregnated with colouring matter. The annexed figure will give an idea of the internal structure, which is manifest even in apparently simple dense wood, when viewed under a strong magnifying glass. It represents the transverse section of a pine bough, the portion marked *a* representing young ligneous cells, those marked *b* the matured cells.

Fig. 92.



Most plants contain in the inner and outer bark a substance which has an astringent taste, is soluble in water, and which is known by the name of *tannin*.

Linen is the inner bark of the flax plant. During the process of *retting*, the outer bark, by the long-continued influence of moisture and air, passes into decay, and then, after rapid drying, may be rubbed off by bending it backwards and forwards (*breaking*); but the filaments of the inner bark, which do not so readily decay, remain behind, and after being parted into their finest fibrils, and arranged parallel by the so-called *heckling*, form the well-known *flax*. The *tow*, which falls off during this process, consists of tangled fibres of the inner bark.

Flax has a grey colour, because it contains a grey colouring matter, which is not soluble in water and lye, though it becomes soluble in lye by exposing the flax, the thread spun, or the linen woven from it, during a long time, to the action of light, water, and air. This is done in the bleaching-yard by spreading it on the grass (*grass bleaching*). Bleaching may be accomplished more rapidly by the use of chloride of lime.

Bast.—Soak the bark of the lime-tree in water till the outer bark is decomposed, and has become brittle; when

it is dry, the inner fibrous part of the bark can be peeled from it, and it then forms the *lime bast*, used for tying up plants. The outer covering of the trees, which is commonly, but erroneously, called bark, consists by no means of the proper bark only, but of two essentially different parts, which have grown very closely together; the external layer is the true bark, the inner is the bast.

Cotton consists of delicate hollow hairs, which form in the cotton-plant in considerable quantities around the seeds. As it exists in nature it is beautifully white. It is nearly pure cellulin, whereas most kinds of vegetable tissue contain other compounds in considerable quantity.

Paper, which is made by beating vegetable fibres, such as those of linen and cotton, with water, and straining the pulp with wire sieves, is of course nearly pure cellulin.

Vegetable Tissue and Water.—*Experiment 1.*—Pour some lukewarm water over sawdust, and let it stand for a day; then squeeze out the liquid through a cloth and boil it; a slight turbidity will appear, and on longer standing a loose sediment will be deposited. Water does not dissolve the woody fibre, though it does the sap contained in it; in this sap, as in that of all other plants, there is always found a substance in solution, which is very analogous to the white of eggs, and which, like it, coagulates in boiling; it is called *vegetable albumin* (see next chapter). There are also contained in the liquid, separated from the albumin, various other substances in solution (mucus, gum, tannin, &c.), which are not precipitated by boiling. If the sawdust, after it has been dried, is treated with alcohol, this will also dissolve some substances (resin, &c.); and so also will ether, lye, and other liquids. Therefore, in the preparation of perfectly pure woody tissue, it must be treated with various solvents in order to remove all the constituents of the sap.

Properties of Cellulin.—Cellulin is insoluble in all ordinary solvents, except those which alter it, but is dissolved by ammonio-hydrate of copper, from which it can be reprecipitated by the addition of hydrochloric acid. Strong sulphuric and nitric acids act upon it in a very remarkable manner.

Vegetable Parchment.—*Experiment 2.*—If a sheet of white

blotting-paper is immersed for a few seconds in a cold mixture of two measures of sulphuric acid with one measure of water, and is then washed very thoroughly (for the last time with very dilute ammonia) and dried, it is changed into a tough substance very like parchment, which does not appear to differ from the original paper in composition. It is now much used for industrial purposes.

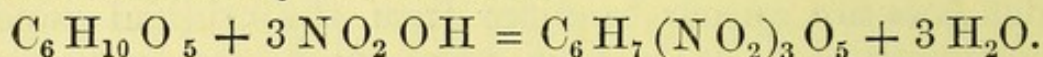
Experiment 3.—If cellulin is beaten for some time in a mortar with strong sulphuric acid, it becomes liquid, and is converted into dextrin and after adding water and boiling into glucose. The excess of acid may be removed by chalk, and the filtered liquid on evaporation will yield both compounds.

Gun-cotton.—*Experiment 4.*—Mix one ounce of the strongest nitric acid (sp. gr. 1.5) with three ounces of strong sulphuric acid; pour the mixture into a porcelain mortar, or a cup, and press into it with the pestle as much cotton (cotton-wool, calico, printing-paper, &c.) as can be moistened by the acid. When the cotton has soaked for twenty-four hours, it is taken out with a glass rod, put into a vessel containing water, and washed repeatedly with fresh quantities of water, until it no longer reddens blue test-paper. The cotton is then squeezed out with the hand, spread upon a sheet of paper, and dried in an airy place. It is dangerous to dry it upon a stove, as it easily takes fire.

If the *gun-cotton* thus prepared is struck smartly with a hammer upon an iron anvil, it detonates violently. This is a dangerous experiment. When touched with a hot wire or a lighted match, it burns instantaneously, without leaving any residue; when fire-arms are loaded with it, it acts like gunpowder, but its explosive power is much greater than that of the latter. Gun-cotton being, therefore, an exceedingly dangerous substance, the *greatest caution* is indispensable in conducting experiments with it, and only very small quantities should be used at once. Another kind of gun-cotton dissolves in a mixture of alcohol and ether into a syrupy liquid, which on spontaneous evaporation leaves the cotton behind in the form of a transparent film. This solution is called *collodion*.

It is used instead of court-plaster, and for making small air-balloons, &c.

The chemical change which takes place in the manufacture of gun-cotton has already been described (page 305). It is described by the formula:



It appears to be a nitrate and not a nitro-compound.

The use of the sulphuric acid is to absorb the water which is formed in the reaction, and thus keep the nitric acid at its maximum strength. The explosive force of gun-cotton is due to the large quantity of oxygen which it contains, and to the fact that when burnt it is entirely converted into gases, which occupy an enormously larger volume than the substance itself.

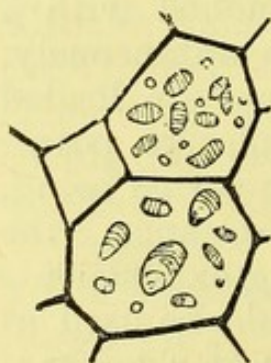
Experiment 5.—Scraps of well-prepared gun-cotton may be ignited on the palm of the hand without danger, in consequence of the suddenness with which the change to gas is effected. Spread a little gunpowder on a plate, and rest a few grains of gun-cotton on it. The gun-cotton can be inflamed without igniting the gunpowder.

STARCH.

A mealy substance, which is known under the name of *starch*, or *fecula*, is deposited in most vegetables, particularly at the period of ripening, from the juices with which the cells of the plants are filled.

It appears to the naked eye like particles of meal, but

Fig. 93.



under a powerful microscope it is found to consist of small, generally regular grains or globules. Their position in the plant is shown in the annexed figure, which represents a section of some of the cells of a potato. Potatoes, grain, and leguminous plants are very rich in starch.

Potatoes.—*Experiment 1.*—Rasp some potatoes on a grater, knead the pulp thus obtained with water, and squeeze it in a linen cloth; the *fibrous particles* of the cells remain behind, but the juice, together with a large portion of the starch, runs through. If you let the turbid liquid remain

quiet for some hours, it becomes clear, because the heavier *starch* settles at the bottom. Now decant the liquid, wash the starch several times with fresh water, allowing it to settle each time, and then dry it in a moderately warm place.

Experiment 2.—Heat in a flask the clear liquid decanted from the starch; it becomes turbid when the heat approaches the boiling point, and, after boiling for a few moments, deposits a flaky, greyish-white substance, which is to be collected on a filter. It is the same substance already referred to (page 310), *vegetable albumin*, characterised by its property of dissolving in cold and warm water, but of coagulating in boiling water. It contains nitrogen, which the starch does not.

Experiment 3.—Put some of the coagulated albumin upon a piece of platinum foil, and heat it over a lamp; it will burn and emit a very disagreeable empyreumatic odour. When starch is treated in the same manner, it also gives off an empyreumatic, but far less unpleasant smell. Many nitrogenous substances comport themselves in this respect like albumin; many non-nitrogenous substances, like starch; therefore, when a piece of woollen cloth is singed, it diffuses a far more disagreeable odour than a piece of cotton or linen, because nitrogen is contained in the wool, but not in the cotton or linen.

Experiment 4.—Mix twenty drops of sulphuric acid with three ounces of water, and pour this acid water upon a potato cut in thin slices; after standing twenty-four hours the slices are to be taken out, washed with water till they no longer redden litmus and then dried. During this process the potatoes lose their juices, and also their albumin and colouring matter, and after drying, form a solid, mealy, white, and tasteless substance, which swells up and becomes soft when boiling water is poured upon it. Potatoes dried without this treatment become grey and horn-like, and acquire an unpleasant smell.

Peas.—*Experiment 5.*—Pour a handful of split peas into a capacious vessel containing water, and let it stand for some days in a warm room; a great deal of water is absorbed by the peas, causing them to swell up, and finally to become so soft that they can easily be mashed between

the fingers. When in this state, bruise them in a mortar, and add sufficient water to form with them a thin paste, which may be squeezed out by means of a linen cloth. Here, also, we obtain, as from potatoes—1, a *fibrous substance*, which remains on the cloth; 2, starch, which is deposited, after standing, from the turbid liquid; 3, *vegetable albumin*, when the decanted liquid is heated to boiling.

Experiment 6.—When you have separated the vegetable albumin, by boiling and filtering, from the above-mentioned liquid, add to the latter a few drops of any acid; a flaky white body will again be deposited; this is called *legumin*, or *vegetable casein* (cheesy matter), on account of its great similarity to the cheese contained in milk (animal casein) in its composition and also in its properties. Vegetable casein, like vegetable albumin, contains *nitrogen*; but it is distinguished from the latter by this, namely, that it is not coagulated by boiling, though it is *by acids*. It occurs in the juice of many plants, but it is most abundant in the seeds of leguminous plants; potatoes, likewise, contain small quantities of it.

Wheat-flour.—*Experiment 7.*—Moisten a handful of wheat-flour with sufficient water to form a stiff paste when triturated in a mortar; enclose it in a piece of thick linen, and knead it frequently under water as long as the liquid which runs through continues to have a milky appearance. After standing some time, a white powder will settle from the turbid water; this is *wheat starch*.

Starch is one of the principal constituents of flour, as, indeed, of all sorts of meal; the second constituent remains behind in the cloth, mixed with vegetable fibre, and is a viscous, tough, grey substance, which has received the name *gluten*. The gluten only swells up in water, without being completely dissolved; in its constitution it corresponds exactly with albumin, and, like this, contains *nitrogen*.

When the water decanted from the starch is boiled, it becomes turbid, and when partially evaporated, yields a flocculent precipitate; thus wheat-flour contains also some *vegetable albumin*.

If the results of these experiments are grouped together, we shall find that there are always present in potatoes and

peas, and also in wheat-flour, the two non-nitrogenous substances, vegetable tissue and starch, and also one or several of the nitrogenous compounds, vegetable albumin, casein, and fibrin.

These nitrogenous substances are often known as *albuminoid compounds*. Small quantities of them occur in the sap of every plant. We shall learn more of them in the next chapter.

Potato starch exhibits, under the microscope, the form of egg-shaped grains, consisting of many scales overlapping each other; it glistens in the sun, is hard to the touch, and has always more of a pulverulent than of a concrete character.

Fig. 94.



Fig. 95.

In the *starch of peas* many of the grains are concave longitudinally, while others seem to be formed by the growing together of several globules.



Fig. 96.

Wheat starch consists of dull, flattened, lenticular, grains, which, when moist, readily adhere to each other, on which account the wheat starch of commerce always comes in loose lumps.



Arrowroot is a starchy meal, which is prepared in the East and West Indies from the roots of some marsh plants.

Experiment 8.—Heat in a vessel half a drachm of starch, with an ounce or an ounce and a half of water, constantly stirring it till it boils; the mixture first becomes slimy, and finally as thick as a jelly. The grains of starch absorb water, and swell up, so that the single membranes break open, though starch is not soluble in water. This swollen starch is well known for its adhesive properties, and is variously employed as a means of thickening printing colours. When linen and other woven fabrics are passed through a thin paste of starch, they acquire, after

drying, a degree of stiffness, and by ironing or strong rubbing and pressing, a bright gloss (starching). The swelling of many of our most common articles of food, such as rice, groats, barley, beans, peas, lentils, &c., when boiled with water, is now readily explained by their containing a large quantity of starch.

Experiment 9.—Dilute some starch paste with a large proportion of water, and add to it a few drops of tincture of iodine; an intensely deep *blue* liquid (*iodide of starch*) is produced. The same colour may be perceived by dropping some tincture of iodine upon flour, potatoes, carrots, &c. We have in iodine an extremely delicate test for starch. Iodine compounds do not give this reaction, until the iodine is liberated. When the blue compound is gently heated, the colour is discharged, but reappears on cooling. The reason of this is unknown.

DEXTRIN, OR BRITISH GUM.

Experiment 1.—If starch is heated in a ladle over a gentle flame, and during the heating (roasting) is constantly stirred to prevent its burning and baking on the bottom of the ladle, it acquires after a while a yellow colour, and then possesses the new property of dissolving, both in cold and in hot water, into a mucilaginous liquid. (Common starch is entirely insoluble in cold water, and only swells up in hot water.) Starch thus transformed is called *dextrin*, *starch gum*, or *British gum*. It is well adapted for the thickening of colours and mordants in calico-printing, and therefore is now prepared on an extensive scale, usually by roasting starch. It is not turned blue by iodine as starch is.

Experiment 2.—Mix thoroughly, in a small dish, half an ounce of starch with one drachm of water and four drops of nitric acid; let the mixture dry in the air, and then place it on the plate of a heated oven, which is just hot enough to hiss feebly when touched with the moistened finger. After some hours, all the nitric acid will be expelled, and the starch will dissolve almost entirely in cold water, and completely in hot water. Dextrin thus made is white, or has only a slight yellowish tinge.

Experiment 3.—Make a paste of potato starch by boiling

starch with water, and while still hot, add to it in a saucer, some drops of sulphuric acid with constant stirring. That this acid effects a change is evident, for the viscid mass very soon becomes a thin liquid. Now place the saucer on a jar, in which some water is simmering (steam-bath), and let it remain over the hot steam (the contents of the saucer not being heated quite to the boiling point) until the liquid has become semi-transparent. When this is the case, add prepared chalk by small portions at a time to the liquid, until it ceases to give an acid reaction, and after having filtered it from the gypsum, leave it to evaporate in a warm place. The dry residue has an amorphous, vitreous appearance and little taste, and dissolves in water, forming a transparent viscid fluid. It consists of dextrin and sugar; and if strong alcohol be added the dextrin is thrown down.

Fig. 97.



Glucose, or Starch-sugar.—*Experiment 4.*—Repeat the former experiment, with the following deviation. Bring to brisk boiling two ounces and a half of water, to which twenty drops of sulphuric acid have been added, and then add one ounce of starch mixed with a little water, forming a paste, but only in small quantities at once, that the boiling may not be interrupted. When all the starch has disappeared, let the mixture boil for some time, then neutralise the acid with chalk, and evaporate the filtered liquid to the consistency of a thick syrup. It possesses a very sweet taste, and consists mainly of a solution of sugar in water. The *starch-syrup* thus made, as well as the white, solid *starch-sugar*, easily prepared from it, are now both articles of commerce.

Starch, as shown by these experiments, is converted by sulphuric acid, or moderate heating, into dextrin and sugar; on more prolonged heating, into sugar. In the latter case, also, dextrin is first formed, but this soon passes into sugar.

It has not yet been explained how this effect is produced.

Malt and Diastase.—*Experiment 5.*—Pour two ounces of lukewarm water upon a quarter of an ounce of coarsely

pulverised barley-malt; let the mixture remain some hours near a fire or stove, or in the sun, and then strain it through a linen cloth; there is found in the filtrate a substance not yet well known, called *diastase*, by means of which the starch may be converted into dextrin and sugar in the same way as by sulphuric acid.

Experiment 6.—Rub a quarter part of the diastase with some hot starch paste, made of a quarter of an ounce of potato starch and two ounces of water; heat the mixture moderately, but not above 149° F. (65° C.), until the paste is formed into a thin, transparent liquid. Now strain through a cloth, and let the liquid evaporate in a warm place. The mass remaining behind contains a soluble variety of starch, *dextrin* and *maltose*, the latter a kind of sugar.

Experiment 7.—Treat the other three-quarters of the diastase in the same way, but prolong the heating for several hours, which may be most conveniently done on the hearth of a stove or fireplace, applying to the liquid a heat not above 158° F. (70° C.) or 167° F. (75° C.). Here dextrin and maltose are first formed; but the dextrin is converted on further heating into maltose, as may easily be perceived by its taste.

The remarkable change which malt causes in starch is to be ascribed to the *diastase* contained in the malt. The mode in which the diastase acts is, however, unknown. At 212° F. (100° C.), consequently, on boiling the liquid, the effect of the malt (diastase) is destroyed. The process of forming sugar by means of the diastase of malt is of great importance to the brewer and distiller, as in the manufacture of beer from barley or wheat, or spirit from rye and potatoes, the starch of these substances must always be previously converted into sugar and dextrin, before fermentation and the consequent formation of alcohol can take place. Beer always contains dextrin, as will be seen if strong alcohol is added to it.

The taste of *malt* is sweet and mucilaginous, because the conversion of the starch into dextrin and sugar commences during germination, the further progress of which is arrested in this case by drying. If the germinated barley is allowed to continue growing, as it does

in the open fields, all the starch gradually vanishes from the grain, and passes, in the form of dextrin and sugar, into the juice of the young plant, as is obvious from the sweet taste of the latter, and from its mucilaginous feeling when rubbed between the fingers.

A similar metamorphosis is also clearly to be perceived in potatoes, and in fruits during their ripening.

GUMS.

From the stem and branches of some trees, such as the cherry and plum trees, and some species of acacia, a transparent viscid liquid exudes, which hardens into globular masses. These exudations are called natural gums. There are several kinds; but the most important are gum arabic, gum senegal, and gum tragacanth. The characteristic ingredients of these gums are closely related in composition and properties to the starches and sugars.

Gum Arabic.—The best known of these gums is gum arabic, which exudes spontaneously from several species of acacia in Africa. The finer sorts of it are white; the more common kinds have a yellow or brown colour. When well dried, it is so hard and brittle that it may be reduced to a powder by pounding. It is soluble in water.

Experiment 1.—Pour two drachms of cold water on one drachm of gum arabic, and occasionally stir the mixture; the gum will, after a few days, entirely *dissolve* in the water, forming a viscid transparent *mucilage*, which may be diluted at pleasure with more water. This mucilage has great adhesiveness, for which reason it is often used, instead of paste or glue, for joining together paper, &c., or for converting a pulverulent substance into a coherent mass (crayons, pastilles, &c.); it has, moreover, a thick consistency, and hence is variously employed in calico-printing as a thickening material for colours and mordants, and in finishing and dressing operations. A variety of gum obtained from the shores of the Senegal, whence it has been called *gum senegal*, is peculiarly well adapted for the latter purpose, as it yields a thicker mucilage than the common gum arabic.

Experiment 2.—Pour some drops of mucilage of gum

arabic into alcohol; they will not mix with each other, as the gum is insoluble in alcohol. If the mucilage is previously mixed with water and a little dilute hydrochloric acid, and is then added to the alcohol, a turbidity ensues, and afterwards a flaky precipitate will subside; accordingly, alcohol may be used for removing gum from those liquids which contain it. The precipitate consists of the substance called *arabin*, or *arabic acid*, which in the gum is combined with lime, magnesia, and potash.

Gum tragacanth is also a vegetable exudation, well-known as a stiffening material, and as forming with water an adhesive paste; it exudes from a shrub which grows in Greece and Turkey, and it occurs in commerce in the form of white, tortuous filaments, or bands. It contains arabin and another compound called *bassorin*.

SUGARS.

Glucose.—We have already seen how cellulin, starch, and dextrin may be converted into a kind of sugar. We shall afterwards find that ordinary sugar, or sucrose, can be changed into glucose with even greater ease. There are two varieties of glucose, which, from the direction in which their solutions deflect a ray of polarised light, are called *dextro-glucose* and *levo-glucose*. They are, however, identical in composition, and nearly identical in chemical properties, so that their separate study need not occupy us.

Glucose is formed also in many vegetables, and is especially abundant in fruits; as, for example, in plums, cherries, pears, figs, grapes, &c. Honey and the white, sweet grains in raisins, consist of it. On account of this origin, glucose is sometimes called *grape-sugar*.

If you taste a dried granule of sugar from a raisin, and then a little common sugar, you will at once perceive that the former is much *less sweet* than the latter; one ounce of common sugar has the same sweetening power as two ounces and a half of grape-sugar. The solubility of these two varieties of sugar in water is likewise very different, grape-sugar dissolving in it much less readily and more slowly than common sugar.

Sucrose or Cane-sugar.—Is found naturally in the sugar-

cane, beet-root, maple, and many fruits. It has not as yet been prepared by any artificial process.

The operations whereby cane-sugar is obtained on a large scale are the following :

1. *Expressing* the juice from the sugar-cane or the rasped bulb of the beet, either by strongly squeezing, or by hydrostatic pressure.

2. *Boiling down* the juice with the addition of lime, by which several foreign substances are precipitated, until it acquires the consistency of a thick syrup; on cooling, the crude sugar is deposited from it, in brownish-yellow crystalline grains (*raw sugar*, or *Muscovado sugar*). The liquid sugar which does not crystallize is allowed to drain off, and forms the well-known treacle (*molasses*).

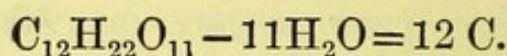
3. *Refining* the raw sugar—that is, the removal of the brown syrup still adhering to it. This is done—*a*), by redissolving the raw sugar in a little water and filtering; *b*), by a second filtration through coarsely-ground animal charcoal, which retains the colouring matter; *c*), by evaporating the clarified solution in vacuum pans. The concentrated syrup is then allowed to cool in moulds of a conical shape, stirring it frequently to disturb the crystallization. A solid mass, consisting of small fragmentary crystals, the common *loaf-sugar*, is obtained, from which the remaining liquid sugar is removed by causing a concentrated solution of crystallizable sugar to percolate gradually through. The thoroughly purified and glistening white sugar is called refined loaf-sugar; that which is not so completely clarified, and has a yellowish tinge, is the common loaf-sugar.

Experiment 1.—Dissolve half an ounce of sugar in a quarter of an ounce of hot water; the viscous solution is called *white (simple) syrup*. If this solution is put in a cup, and set aside in a warm place, and evaporated slowly, the sugar will separate from it, crystallizing in oblique prisms. In a similar manner *white sugar candy* is prepared on a large scale from refined sugar, brown candy from raw sugar. As the crystals deposit more readily on substances having a rough than on those having a smooth surface, fine threads or pieces of wood are stretched

across the vessels containing the syrup, and they soon become coated with crystals.

Cane-sugar, as already stated, has a much sweeter taste than grape-sugar; therefore, when used as a sweetening agent, it possesses a far greater value than the latter.

Experiment 2.—Put two or three lumps of sugar in a breakfast cup standing in a basin, moisten them with boiling water, and then add about twice the bulk of strong sulphuric acid. The acid removes the elements of water from the sugar, the mixture swells up enormously, and a black mass of charcoal remains:



Experiment 3.—The two sorts of sugar may be accurately distinguished from each other by the *copper test*. First add, in test-tubes, to very dilute solutions of each kind of sugar, some drops of a solution of copper sulphate then excess of solution of potash. No turbidity occurs in either case, as sugar prevents the precipitation of copper as cupric hydrate. Place both tubes in boiling water; the liquid containing the grape-sugar assumes in a few minutes a reddish-yellow colour, while that containing the cane-sugar remains blue. The grape-sugar is able to abstract oxygen from the cupric salt, whereby reddish-yellow cuprous oxide is formed. Cane-sugar, when long boiled with the copper test, will also reduce the copper, but only because it first passes into glucose. Grape-sugar may be distinctly recognised by this test, even in an extremely diluted solution.

Fehling's test. Experiment 4.—Dissolve 30 grains of crystallised cupric sulphate in a little water. In another portion of water dissolve $3\frac{1}{2}$ ounces of Rochelle salt (tartrate of potassium and sodium) and $2\frac{1}{2}$ ounces of caustic potash. Mix the solutions and dilute with water to one pint.

The bright blue solution so obtained is excellent for sugar testing. It may be boiled without change, but on the addition to the boiling liquid of a dilute solution of glucose, maltose, or lactose a red precipitate of Cu_2O is at once formed. One-third ounce of the solution would be decolorized by about 2 grains of glucose.

Experiment 5.—Boil in a dish half an ounce of sugar with one drachm of water, until the viscous solution begins to assume a yellowish tinge; then pour it upon a plate previously smeared with a little olive oil. The transparent brittle mass is melted sugar in an *amorphous* state (barley-sugar). The sugar is first dissolved by the water; on boiling, the water is again evaporated, and the sugar passes gradually from the dissolved to the melted state. The yellowish colour indicates that all the water has passed off and that the sugar is on the point of becoming burnt.

If the transparent sugar is kept for some weeks, it becomes opaque and crystalline, returning to its original condition.

Experiment 6.—Repeat the former experiment, but without stopping the heating on the appearance of the yellow colour; the sugar will grow darker, until it finally attains a brownish-black colour, and exhales, at the same time, a peculiar empyreumatic odour. On cooling, it is obtained as a hard, almost black mass, which soon deliquesces in the air, forming a dark syrup, which is called *caramel* or *burnt sugar*. A couple of drops of it impart to a large vessel filled with water the appearance of Jamaica rum. On account of its strong colouring properties, burnt sugar is much used for imparting to liquors—vinegar, alcohol, &c.—a yellow or brown colour.

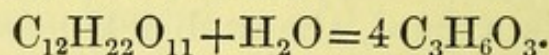
Experiment 7.—When exposed to a still stronger heat, the sugar becomes charred, and finally burns like wood, as may easily be seen by holding a piece of it on platinum foil over an alcohol flame. The flame indicates also that inflammable gases are evolved. Pure sugar leaves scarcely any residue. If it contains lime, white ashes remain behind upon the foil, which do not volatilise even at the strongest heat.

Experiment 8.—If you add a few drops of dilute sulphuric acid to a dilute boiling solution of sugar, and continue to boil for some time, a liquid is obtained which does not crystallize on evaporation. If you now apply the copper test, you will find that it contains grape-sugar; cane-sugar is therefore converted by boiling with sulphuric acid into grape-sugar. This metamorphosis is

produced also by fermentation. If sugar is boiled for a long time with diluted sulphuric acid, it finally passes into a brown substance, resembling humus. When boiled with nitric or other acids, which yield oxygen, it oxidises first into saccharic acid, then into oxalic acid, and finally into carbonic acid and water.

Sugar, as though it were an acid, can combine in fixed proportions with oxide of lead, lime, and many other bases; but it thereby loses its sweet taste.

Lactose, or Milk-sugar, is that particular kind of sugar which occurs in milk, and imparts to it its agreeable sweetish taste. It is obtained in hard, white, crystalline masses, by evaporating the sweet whey. Sugar of milk is much less sweet to the taste than grape-sugar, and requires six parts of cold water for its solution. Like glucose it reduces copper. It is well known that milk becomes sour by standing for some days; this is owing to the sugar of milk being gradually converted into a peculiar acid, called lactic acid.



Mannite, $\text{C}_6\text{H}_{14}\text{O}_6$, is a substance resembling sugar, constituting the principal part of manna (the concrete sweet juice of some species of the ash, growing principally in Italy).

Although the composition of the preceding compounds cannot yet be said to be established, it appears highly probable that they are closely related to alcohols of high valency. It is indeed known that mannite, $\text{C}_6\text{H}_{14}\text{O}_6$, is a true hexatomic alcohol, and that its formula should be written $\text{C}_6\text{H}_8(\text{OH})_6$. The fact that glucose, by combining with hydrogen, yields mannite, would seem to indicate that it stands to the latter compound in the same relation that aldehyde does to alcohol:

Alcohol $\text{C}_2\text{H}_6\text{O}$.
Aldehyde $\text{C}_2\text{H}_4\text{O}$.

Mannite $\text{C}_6\text{H}_{14}\text{O}_6$.
Glucose $\text{C}_6\text{H}_{12}\text{O}_6$.

CHAPTER IV.

ALBUMINOID GROUP.

THE members belonging to this group are among the most complex and least understood of compounds. They are found both in the animal and vegetal kingdoms, and are essentially necessary to the processes of life. No animal can live unless its food contains one or more of them, and they are, therefore, sometimes spoken of as the nitrogenous, or *flesh-forming* elements of food. They contain five, or even six elements, and they resemble one another so closely in composition as to suggest the idea that they are all modifications of some one substance. But no satisfactory formulæ can be deduced for them, and their composition must, therefore, be expressed in parts per cent. The following analysis of albumin may be given as an example. The simplest formula which can be deduced from it is $C_{72}H_{112}N_{18}SO_{22}$.

Carbon	53·3
Hydrogen	7·1
Nitrogen	15·7
Sulphur	1·8
Oxygen	22·1
	100·0

The following are the most important members of this group :

Albumin	}	Derived both from animals and vegetables.
Fibrin		
Casein, or Legumin	}	Derived only from animals.
Ossein, or bone cartilage		
Gelatin		

The first three, as found in the vegetable kingdom, have been described to some extent in the last chapter. *Gluten*, obtained by washing wheat flour (page 314), appears to be a mixture of substances like casein and fibrin. When it is boiled with alcohol the former dissolves, and can be obtained by evaporation.

Albumin.—The *white* of eggs consists of cells, in which is contained a colourless alkaline liquid, the *albumin*. On evaporation, we obtain from it one-eighth of solid albumin, the rest is water. When burnt, it leaves behind common salt, carbonate, phosphate and sulphate of sodium, and phosphate of calcium. That albumin, when briskly beaten up, yields a porous light froth, that it becomes insoluble and coagulates by heating, &c., are well-known facts. On account of the latter property, it is used for clarifying turbid liquids, especially saccharine juices.

Experiment 1.—Stir up some honey in warm water, add a little albumin to the turbid solution obtained, and heat the mixture to boiling. The albumin seizes upon the foreign substances floating in the liquid, bears them to the surface, and incloses them within itself as it coagulates; the liquid thereby becomes clear and transparent, and may be separated by a strainer from the coagulated albumin.

Experiment 2.—Add water to the white of a fresh egg, and stir. As the cells break up a solution of albumin is obtained, which after standing may be poured off from the fibrous residue. With this solution several experiments may be made.

1. Boil a little in a test-tube, and observe that it becomes opaque (coagulates).

2. Add to fresh portions of the solution small quantities of dilute nitric acid, strong alcohol and solution of mercuric chloride (corrosive sublimate). In each case the albumin will become insoluble.

3. Other portions may be treated with acetic acid, with common salt, and with many other substances, without any precipitate being produced. It is therefore seen that some acids and salts coagulate albumin while others do not.

The *yolk of eggs* consists of albumin holding in suspension yellow drops of oil. On account of the albumin contained

in it, it coagulates when heated, and the fat (*oil of the yolk of eggs*) may be extracted from it by strong pressure, or by agitation with ether. *Phosphorus* is contained in the oil of yolk of eggs.

We shall soon see that albumin is contained in the blood. It is also an important constituent of the juice of flesh.

Fibrin.—The solid substance of muscular tissue (flesh) consists of this compound, which is also present (or a substance similar to it) in a soluble condition in the blood. From the latter it can readily be obtained, though only in the insoluble form. The blood consists of a clear pale-yellow liquid, the *plasma*, or *liquor sanguinis*, in which float innumerable red discs like pieces of money, which are called the *red corpuscles*. To these the colour of the blood is due.

Experiment 1.—If you let the freshly-drawn blood of an animal remain standing quietly in a vessel, it will, in a short time, undergo a change; it coagulates, forming a dark-red jelly (*clot, coagulum*), which contracts on longer standing, and a yellowish liquid is separated (*serum*). When the latter is heated to boiling, it coagulates to a white jelly; the serum consists mainly of a solution of *albumin*. There are two substances in the coagulum, one of which called *hæmoglobin* dissolves by washing in water, yielding a red solution (*colouring matter of the blood*, the principal constituent of the blood corpuscles), while the other remains behind as a white fibrous mass (*blood fibrin*). Accordingly, the most important proximate constituents of the blood are water, albumin, blood corpuscles, and animal fibrin, which, on the standing of the blood, are transposed in the following manner:

From	Water, Albumin,	Blood Corpuscles, Fibrin,
are formed	<u>Serum</u>	and <u>Coagulum.</u>

It is a distinguishing peculiarity of the colouring matter of the blood, that it always contains iron.

Experiment 2.—If the blood freshly drawn from the veins is beaten up during cooling, it does not coagulate; the fibrin certainly becomes insoluble, but it separates as a thread-like coherent mass, which, when kneaded for

Experiment 2.—If you add to the filtered liquor a few drops of hydrochloric acid, the *casein* separates from it as a white flaky mass; accordingly, the animal casein is coagulated and rendered insoluble by acid, in the same manner as vegetable casein. Pure casein is insoluble in water, but it dissolves in it when alkalies are present.

Experiment 3.—If you filter the casein from the liquid, and then boil the latter, it again becomes turbid, although less so than before. It is the *albumin* which separates, small quantities of it being present in all milk.

Experiment 4.—Let a small piece of the dried membrane of the stomach of a calf (*rennet*) remain standing one night in a small quantity of water, and afterwards mix this water with a quart of new milk; the milk, after having stood for some hours in a warm place, will coagulate into a gelatinous mass, which is to be put upon a filter. What remains behind consists of an intimate mixture of the curdled *casein with globules of fat*. By pressing and drying, we obtain from it the so-called *cream- or new milk-cheese*.

Experiment 5.—Separate the filtered liquid (*sweet whey*) from its albumin by boiling, and, having again filtered it, evaporate until only a few ounces of it remain. If left in a warm place, hard, prismatic white crystals of *sugar of milk* will be deposited (page 324). By this method sugar of milk is procured on a large scale. Consequently the sweet whey is to be regarded principally as a solution of sugar of milk (together with some albumin and some salts) in water.

Experiment 6.—Dissolve again in water the sugar of milk obtained, and put a piece of *rennet* in the solution; the liquid will soon become sour in a warm place, because the sugar of milk is converted into *lactic acid*.

Experiment 7.—The coagulation of the milk, which was produced by the rennet in a few hours, is effected instantaneously by the addition of *acids*, as is rendered obvious by adding a few drops of some acid to warm milk. In this curdled mass are contained all the casein and fatty particles of the milk (*cheese and butter*).

Experiment 8.—Fill a bottle with fresh milk, close it, and keep it inverted, from twenty-four to thirty-six hours, in a cool place; then loosen the stopper a little, so

that the lower, thinner portion of the milk (*blue* or *skim milk*) may run off, but the upper, thicker part (*cream*) remain behind. On standing, the lighter oil-globules of the milk ascend, and form on the surface the well-known fatty, thick cream. If this is shaken for some time, the membranes of the oil-globules are torn, and the latter then unite together, forming masses of *butter*. The thin milk which passes off from beneath may be separated, in the way already described, into casein, albumin, and sugar of milk.

Experiment 9.—If you let milk stand for some time in *open vessels*, its sugar of milk is gradually converted into *lactic acid* (p. 324), and this, like every other acid, causes a curdling of the milk, and, at the same time, its well-known sour taste. But the curdling first commences after most of the oil-globules have collected on the surface (*sour cream*). From this cream butter is often prepared, and therefore the buttermilk remaining (a mixture of curdled casein, lactic acid, and water, with some particles of butter remaining behind) has an acid taste. The so-called *curd* beneath the cream now contains only traces of fat, and consists accordingly of water, lactic acid, and coagulated casein. By pressing we obtain from it the *sour whey*, and, as a residuum, the coagulated casein, from which common *skim-milk cheese* is made. When kept damp this undergoes a kind of putrefaction, and a soft saponaceous mass is produced. If the putrefaction advances still farther, there will be finally generated also volatile compounds of strong smell (ordinary cheese).

Gelatin.—The organic matter (about one-third) of bones consists of a compound called *ossein*, or *bone cartilage*, which is insoluble in water, but which, after prolonged boiling, changes into the well-known substance gelatin, which, though scarcely soluble in cold water, dissolves easily in hot. The skin and some other parts of the animal body experience the same change when long boiled with water. If the water is made to boil at a higher temperature (about 220° F., or 105° C.), by enclosing it in a strong iron vessel, the conversion is effected much more rapidly, as it is also by superheated steam. *Isinglass*, which is prepared from the air-bladder of the sturgeon, is nearly pure

CHAPTER V.

FERMENTATION.—ETHYL COMPOUNDS.

THE compounds described in the last chapter, when exposed to the air in a moist state and between certain limits of temperature, experience a peculiar change, which is called *putrefaction*. They decompose and are converted into simpler compounds, some of which are gaseous, and possess a foetid odour. This change, which is an example of many similar ones which are known, is due to the absorption from the air of minute germs—seeds, or eggs—of low forms of life. If air be entirely excluded, the change cannot take place, because the germs cannot enter; but when once commenced it will continue without any assistance from the air. The organisms develop and reproduce themselves; they feed on the compounds with which they are in contact, and throw out as excretions the compounds which we know as products of putrefaction.

It is certain that very minute organised germs of many kinds float constantly in the atmosphere, ready to develop and increase whenever they meet with suitable food and suitable conditions of temperature, moisture, &c. Some have been detected, and their developments studied with great care, but even the existence of many of them can only be inferred from their effects.

FERMENTATION.

The albuminoid substances, when in a state of incipient putrefaction, have a remarkable power of causing decompositions in solutions of sugar. These changes, which are akin to putrefaction in their nature, are effected by peculiar organisms, the germs of which are derived from the air.

The organisms feed on the albuminoid substances and sugar jointly, and thereby convert them into new and simpler compounds, which are distinguished from the products of putrefaction by being, for the most part, destitute of offensive odour. Changes of this kind are called *fermentations*, and the substances which produce them *ferments*. There are several kinds of fermentation, but the most important is that known as *vinous*, or *alcoholic* fermentation.

Experiment 1.—Half an ounce of honey (nearly pure glucose) is dissolved in four ounces of water and some gluten or flour, which has been kept in a warm room for about a week till it has begun to putrefy, is added to it; the liquid is then put in a moderately warm place, from 64° F. (18° C.) to 75° F. (24° C.), when it soon enters into fermentation, with the

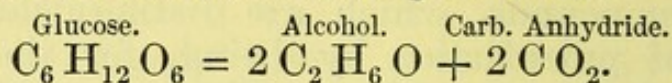
! Fig. 98.



evolution of a large quantity of gas. If you perform the experiment in a bottle furnished with a bent glass tube, one end of which is passed under a second bottle, filled with water, which is inverted over the pneumatic trough, the gas may easily be collected; it consists of carbonic anhydride. If the liquid still retains a sweet taste after the evolution of the gas has ceased, add another portion of the gluten to it, whereby the fermentation is again renewed. Finally, all the saccharine taste will have disappeared, and the liquid will have acquired a spirituous flavour. The fermented liquor is called *mead*; instead of grape-sugar it contains *alcohol*, and this is the cause of its intoxicating effect. A portion of the gluten is found at the bottom of the vessel, converted into a brownish residue.

Yeast.—While the foregoing process is going on, the liquid contains a multitude of minute egg-shaped bodies called *yeast-cells*, because yeast is mainly composed of them, which are the agents by which the sugar is converted into alcohol and carbonic anhydride. Yeast is, of course, the most powerful of ferments.

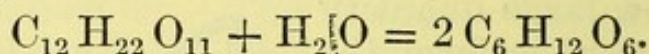
The change which grape-sugar undergoes in fermentation is represented in the following formula :



But some other compounds are also formed in small quantity during the fermentation. Of these the higher alcohols, glycerin, and succinic acid, may be mentioned.

Experiment 2.—Repeat the former experiment, adding, instead of the gluten, a teaspoonful of yeast, or a few fragments of the German dried yeast; the fermentation will now proceed more rapidly and regularly.

Experiment 3.—Instead of honey, take a solution of cane-sugar and add some yeast to it. In this case the fermentation will not begin so soon, for the cane-sugar has first to take up the elements of water and change to grape-sugar :



WINE.

All sweet vegetable juices pass spontaneously into fermentation without the necessity of adding to them a ferment, because they always contain sugar and albuminous substances, in which yeast cells will soon develope.

Experiment 1.—Submit freshly expressed beet-juice to a temperature of from 20° to 25° C. (68° to 77° F.); the juice will soon effervesce, deposit a sediment, and be converted into a spirituous liquid (beet-wine).

In the same manner currant and gooseberry wines are prepared from currants and gooseberries, cider from apples, the so-called cherry-water by fermenting and afterwards distilling cherry-juice, rum by fermenting and afterwards distilling the juice of the sugar-cane, &c.

The most common of all the fermentations of this kind is the fermentation of grape-juice, wine being the result. In order to prepare *white* wine, the grapes are pressed, the juice (*must*) is poured into vats and allowed to remain in them in the cellar, where, as the temperature is tolerably low, the fermentation proceeds so slowly that it is not completed until after some months. The new wine is racked off from the lees, and poured into fresh vats; it still con-

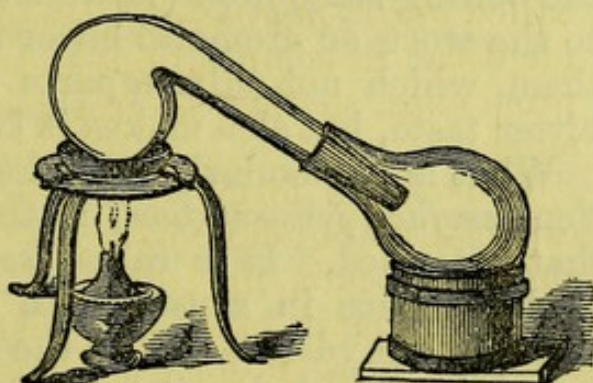
tains a small quantity of sugar and albuminous matter, which are both gradually converted, the former into alcohol and carbonic anhydride, the latter into lees (*secondary fermentation*). In the manufacture of *red wine*, the purple grapes are bruised, and then fermented, together with their skins and stalks; by which means a red colouring matter is extracted from the skins, and tannin from the stalks and seeds, the tannin imparting to this kind of wine its astringent taste. *Sparkling wine* (Champagne) is made by allowing the secondary fermentation to take place in corked-up bottles, whereby the carbonic anhydride formed is retained in the wine.

The acidity of most wines is due to *tartar* or *argol* (acid potassium tartrate, $\text{KHC}_4\text{H}_4\text{O}_6$), which is contained in grapes. A good deal of this subsides when the wine is kept. Minute crystals of it are often seen on the corks of claret bottles.

Many wines, champagne for instance, are sweetened by the addition of sugar, and some (port and sherry) are *fortified*, or rendered stronger, by the addition of spirit.

Experiment 2.—If some wine is put into a retort, and subjected to distillation to a moderate heat, at first the more volatile alcohol, together with certain fragrant ethers, will pass over. A very agreeably smelling spirit is thus obtained, known in commerce under the name of *Cognac*, or *French brandy*. In the wine countries, the lees remaining after the wine is racked off, are generally used for this purpose, since, in the swollen, pap-like state in which they settle in the vats, they retain mechanically a large quantity of wine.

Fig. 99.



BEER.

The manufacture of beer differs essentially from that of wine in this respect: that materials are employed which contain very little sugar already formed, but instead of it

starch, such as barley, wheat, rye, potatoes, &c. Starch cannot, like sugar, be resolved directly into alcohol and carbonic acid; and, when employed in the manufacture of alcohol, it must previously be converted into sugar. This conversion is effected by the diastase of the barley-malt in the so-called mashing process of the brewers and distillers (page 318).

Experiment 1.—Pour a mixture of an ounce and a half of cold water and two ounces of boiling water upon half an ounce of bruised malt, and set it aside for some hours in a warm place, where it will keep a temperature of from 65° to 75° C. (149° to 167° F.); a sweet liquid is thus obtained, composed not only of dextrin and maltose, but containing also the altered gluten, thereby rendered soluble, which was present in the malt. The brewer calls this liquid the *wort*. Strain it with pressure through a cloth, and boil the liquid for some time until it becomes clear and transparent; then let it cool to 86° F. (30° C.), and add to it a teaspoonful of yeast; it will soon begin to ferment, and after some days will clarify again; the clear fermented liquor is *beer*. This is the mode of making the *Berlin white* or *pale beer*, which is not bitter. If during the boiling some hops (flowers of the hop-plant) are added to the wort, an aromatic bitter substance is dissolved from them, which not only imparts to the beer a pleasant and bitter taste, but also makes it keep better.

What is particularly remarkable in the above fermentation (*surface fermentation*) is the great quantity of yeast that is formed. It is called *surface yeast*, it being raised to the surface in consequence of the great evolution of carbonic anhydride, and if the casks are full, it is forced out of the bung-hole; it is the best yeast, and the quantity obtained in the last experiment is sufficient to bring to complete fermentation the wort of a whole pound of malt.

Fig. 100. Its power of exciting fermentation is destroyed when it is rendered quite dry, or when it is boiled, or very finely triturated; and likewise by mixing antiseptic substances with it, as, for instance, alcohol, carbonic acid, sulphurous acid, volatile oils, &c. This yeast, when examined through the microscope, has the form of simple cells (a);



and their increase in the wort takes place in the same manner as in the most simple plants, new cells or buds developing themselves on each globule of the old yeast.

New beer contains, also, some sugar and gluten in solution; therefore, like wine, it undergoes, when kept, a second slight fermentation (*secondary fermentation*). If this is allowed to take place in well-stopped bottles, so that the carbonic anhydride cannot escape, a foaming beer (*bottled beer*) is obtained, in the same way as in the manufacture of sparkling Champagne.

But all the gluten is not separated, even by the second fermentation, and hence the beer undergoes a still further change on being exposed to the air; it is the alcohol, however, which is now altered by the albuminous matter undergoing decomposition; it passes into vinegar, and the beer becomes acid (*acetous fermentation*).

Experiment 2.—Repeat the former experiment, but cool the wort below 50° F. (10° C.) before adding the yeast, and then let the liquid remain in a cool place; a very slow fermentation takes place, which will not be completed for several weeks, perhaps even months. During this process, the carbonic acid is evolved in very small bubbles, and the yeast settles at the *bottom* of the vessel (*bottom fermentation, bottom yeast*). German beer is made in this way.

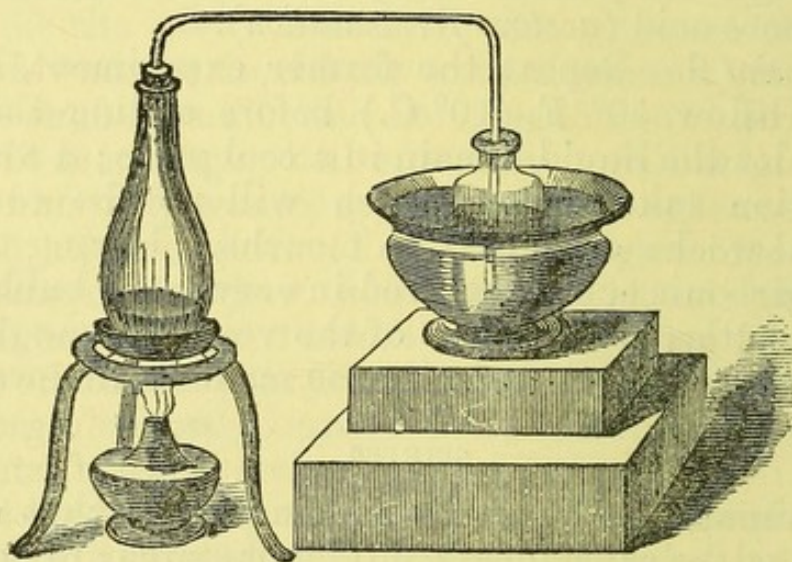
SPIRITS.

The preparation of spirit is similar to that of beer, inasmuch as the substances containing starch are employed in the preparation of it, and as the starch must first be converted into sugar before the fermentation can proceed. This is done, as in the case of beer, by the mashing process—that is, by the operation of the diastase of the malt upon the starch. To this end either boiled and mashed potatoes or rye, or rice, or barley, is mixed with bruised barley-malt and hot water, so as to form a paste, which is to be kept at a temperature of 158° F. (70° C.) until a complete formation of sugar is effected; then brewers' yeast (surface yeast) is added to the sweet mash or wort previously cooled, whereby fermentation is induced. When this fermentation is concluded, the mass is put into a copper still, and the volatile alcohol distilled from the

non-volatile parts (husks, gluten, fibrous matter, &c.). The residue is used as a nourishing food for the fattening of cattle. Formerly simple retorts were used for this distillation, and a weak spirit was obtained, which consisted of about one-third of alcohol and two-thirds of water; but now a more complicated apparatus is employed, by means of which a spirit of double the strength is obtained (*rectified spirit*). The principle upon which this apparatus depends will be explained in the following experiments.

Rectification or strengthening of Spirit.—*Experiment 1.*—Pour three ounces of common spirit of wine into a capacious flask, and carefully distil half of it into a vessel, which is cooled by means of very cold water, or what is still

Fig. 101.

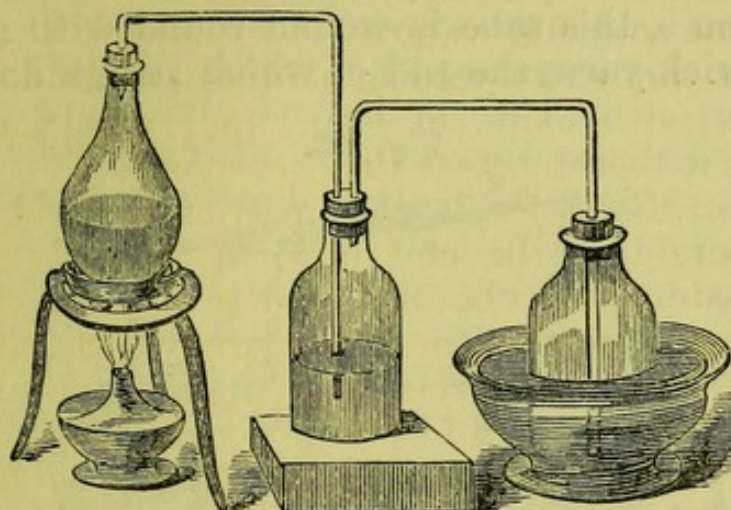


better, by ice. If the spirit contained thirty per cent. of alcohol, then the ounce and a half of alcohol first passing over will contain at least fifty per cent. Alcohol is *more volatile* than water; therefore it passes over first, in company with a smaller quantity of the latter, while the greater part of the water, together with the fusel oil (amyl alcohol, &c.), which might have been contained in the spirit, remains behind in the flask (*phlegm*).

Experiment 2.—If you connect with the flask and the receiver an intermediate vessel, a wide-mouthed bottle, for instance, which is easily done by means of two glass tubes bent at right angles, and a cork perforated with two holes, and then repeat the above experiment of dis-

tillation, the alcoholic vapours passing over will first condense in the middle vessel. But as this vessel is not cooled down the liquid condensed in it will finally also

Fig. 102.



boil, and the vapours thus formed will pass over into the receiver surrounded with cold water, and will there be condensed for the second time. In this manner a double distillation (rectification) is effected. The flask contains boiling spirit (at 30° Tralles*); the intermediate vessel, boiling rectified alcohol (at about 50° Tralles). After the termination of the experiment, the first vessel will contain phlegm; the second, weak spirit; and the third, very strong, highly-rectified spirit (of 70° to 80° Tralles). Common gas-pipe may be used with advantage instead of glass in these experiments.

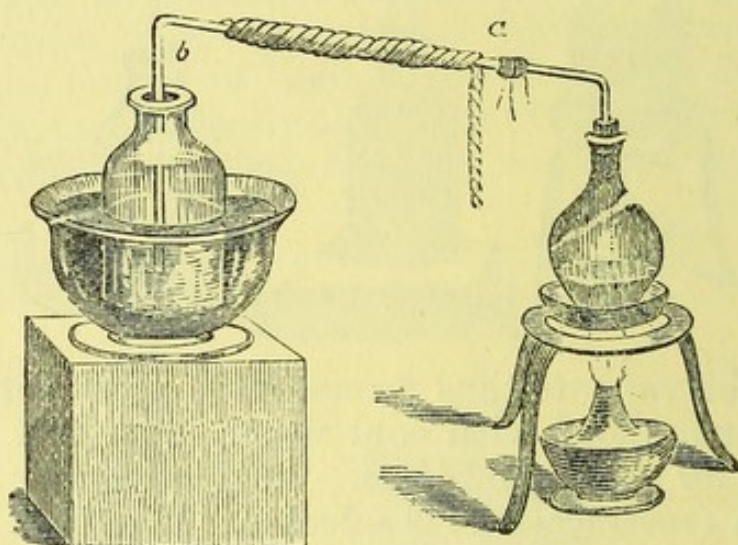
If you adapt to the corks of the first two vessels a couple of thermometers, which shall dip into the liquid, you will find that the liquid in the flask boiled at the commencement of the experiment at 185° F. (85° C.), and at the end of the experiment at from 203° F. (95° C.) to 212° F. (100° C.), while that contained in the second vessel commenced boiling at 176° F. (80° C.), and ended with boiling at from 185° F. (85° C.) to 194° F. (90° C.). It is obvious from this that strong spirit boils at a much lower temperature than that at which weaker spirits

* The alcoholometer of Tralles floats to a figure on the stem, which indicates the percentage of alcohol, *by volume*, at 60° F. (15.5° C.), in the liquor in which it is placed.

boil. The strongest alcohol (absolute) boils at 173° F. (78° C.).

Experiment 3.—Connect with a flask a tolerably large glass tube, which is so bent that its middle part has a slight inclination upwards, as is shown in the annexed figure: from *b*, this tube is wound round with moistened calico, or wick-yarn, the end of which hangs down at *a*.

Fig. 103.



At *a*, bind a strip of cloth (several times folded together and smeared with a few drops of olive-oil) round the tube, so that water from the wick may not run down upon the flask. Now distil as before three ounces of spirit, but during the distillation continually drop cold water upon the wick-yarn, at *b*, in order to cool the vapour of the spirit as it passes over. Catch the water running down the outside of the tube in a vessel placed below the end of the wick-yarn. If the distillation is arrested when about one ounce of the spirit has passed over, we shall have a stronger spirit in the receiver than was obtained in Experiment 1, because, by the partial cooling of the vapour of the spirit, the principal part of the steam was condensed, and, therefore, a vapour richer in alcohol passed into the receiver, while the water condensed in the tube flowed back into the flask.

BREAD.

Experiment 1.—Mix some flour with lukewarm water to

a thick paste, cover it with a board, and let it remain for eight or ten days in a warm place. The paste is gradually altered, and two distinct periods may be observed during the change. In the first place, on the third or fourth day bubbles of air are evolved from it, having a sour, unpleasant smell, and the dough now possesses the capacity of converting sugar into *lactic acid*, as may be readily perceived by adding a little of it to some sugared water, and letting it stand in a warm place. After six or eight days the dough acquires a pleasant smell, and it now acts, when added to a solution of sugar, like yeast; that is, it effects a decomposition of the sugar into *alcohol and carbonic anhydride*. If the dough is allowed to stand still longer, it again acquires an acid taste, but which now proceeds from the *acetic acid* into which the alcohol previously generated is gradually converted (*leaven*). In this state also it excites a spirituous fermentation in sugared water; but this spirituous fermentation is immediately changed into the acid—into *the formation of vinegar*. It is obvious, from what has previously been stated, that the different actions of the flour, when in process of decomposition, upon the sugar, depend upon the albuminous matter, the gluten, contained in the flour; consequently, we might call the slightly altered gluten a *lactic acid* ferment, that which is more altered an *alcohol* ferment, and that which is still further altered a *vinegar* ferment.

Bread.—What thus takes place slowly, proceeds rapidly in the making of bread, since a ferment is purposely added to the flour, which has been stirred up with water.

In the making of white bread, the surface-yeast of beer is used as a ferment. The sugar contained in the flour is thereby resolved into alcohol and carbonic anhydride, which struggle to escape, whereby the tough mass of dough is disintegrated, and rendered light and porous (*rising of the dough*). These substances, together with about half the water employed, volatilise by the rapid heating in an oven having a temperature of from 320° F. (160° C.) to 356° F. (180° C.), and the cellular partitions of the baked bread acquire such solidity, that they retain their form and place even after cooling. But if the heat of the oven is not sufficient, or the dough is

too watery, then the partitions harden too slowly, and, on the escape of the carbonic acid, collapse, or run into each other (*slack baking*). This happens most frequently with dark bread, since, in consequence of its greater amount of gluten, it retains the water more obstinately, and accordingly dries and hardens more slowly than white bread, in which the starch is more abundant.

Leaven is commonly used as a ferment in the preparation of black bread. There is formed, during the process, besides alcohol and carbonic anhydride, a little acetic and lactic acids (perhaps also some butyric acid), which communicate to the bread an acid taste. From three pounds of flour we obtain about four pounds or more of bread: consequently, at least a quarter of the bread consists of water. The light, porous bread dissolves easily in the stomach, we say that it is easily digestible, and that the compact heavy bread is difficultly digestible.

It is known (p. 316) that starch is converted by roasting, into *gum* (dextrin); a part of the starch undergoes, also, this change in the oven, particularly on the surface of the baked bread, which receives the strongest heat from the roof of the oven. If the crust of the hot bread is rubbed over with water, and the bread is then replaced for a few minutes in the oven, some of the dextrin is dissolved, and forms, after the evaporation of the water, the shiny coating which we see on loaves of bread and rolls.

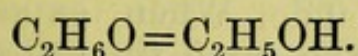
Carbonic anhydride, as applied to the rising of bread, may be more or less advantageously generated in other ways than by the fermentation of sugar; indeed, entirely different substances, which become aeriform on the application of heat, may be used for the purpose.

Experiment 2.—Mix intimately together two grains of finely powdered hydrogen sodium carbonate (bicarbonate of soda), and a drachm and a half of flour, and knead the mixture into a dough with one drachm of water, to which four drops of hydrochloric acid have previously been added. Let the dough remain for some time in a warm place, and then bake it. A porous mass of bread is obtained, because the carbonic anhydride of the sodium salt is expelled by the hydrochloric acid, and raises the dough

while it is still soft. The common salt which is formed remains behind in the bread, and imparts to it a faint saline taste. This method has been introduced in many places for making bread, cakes, &c., on a large scale.

Experiment 3.—Rub a drachm and a half of flour with a few grains of ammonium carbonate, and then knead it with a drachm of lukewarm water into a dough, and treat it as in the last experiment. In this case, also, the mass will become light and porous after the rising and baking, because the salt is rendered aeriform by the heat, and during its escape the particles of the dough are forced asunder. In this way bakers prepare light and spongy cakes, as, for instance, gingerbread, &c. *Aerated* bread is prepared by forcing carbonic anhydride into the dough by powerful pumps. When the pressure is removed, the gas escapes and the dough rises.

ALCOHOL.—ETHYL HYDRATE.



Although alcohol can be prepared by other processes, it is always in practice obtained by the fermentation of sugar. In the preceding pages we have shown how alcohol is formed, how it is rendered stronger, and how it is purified. This is done by partial distillation, or by partial condensation, since the alcohol is more easy to volatilise than water, whilst its vapour is more difficult to condense than steam. But all the water cannot be separated in this way from the alcohol, as the alcohol retains one-tenth part of the water so firmly that it can neither be withdrawn from it by distillation nor by cooling. In order to procure it absolutely anhydrous, a body must be presented to it which has a greater affinity for water, and fixes it so firmly, that it cannot evaporate with the alcohol at the boiling point of the latter. Such a body is quicklime.

Experiment.—Put into a flask one ounce of quicklime that has been broken into small pieces, and pour upon it one ounce of very strong alcohol; connect a receiver with the flask, as in Experiment 1 (p. 338), and let the mix-

ture remain in repose for one day. The lime gradually combines with the water of the alcohol (it slakes), and the latter is procured anhydrous by distilling it off at a moderate heat. The best method of distilling in this case is over the water-bath. *Anhydrous* alcohol is also called *absolute alcohol*. In this experiment, the vessels used must be previously rinsed out, not with water, but with strong alcohol, because the moisture adhering to the vessel would again impart water to the anhydrous alcohol.

Properties of Alcohol.—Alcohol is the hydrate of the monad radical ethyl, C_2H_5 , and has accordingly the formula C_2H_5OH (p. 297). It has a burning taste, and a penetrating, agreeable odour. Strong alcohol, especially absolute alcohol, acts as a poison when swallowed; but when diluted, it is, as is well known, stimulating and intoxicating.

Absolute alcohol only freezes at a very low temperature ($-130^{\circ} C.$); it is therefore well adapted for use in thermometers which are to measure low temperatures. Its specific gravity is 0.8. When exposed to the air it evaporates, and at the same time absorbs water, and so becomes weaker. It may be mixed with water in all proportions. If equal volumes of alcohol and water are mixed, a slight contraction takes place. Alcohol is a very valuable solvent, for it dissolves many substances, resins, for instance, which are insoluble in water. On the other hand, some substances which are soluble in water are insoluble in alcohol.

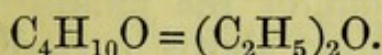
Alcohol, from the small proportion of carbon it contains, burns with a smokeless flame. In burning it is converted into water and carbonic anhydride.

Methylated Spirit.—The Excise permits the sale, duty free, of strong alcohol mixed with ten per cent. of the *wood spirit* (methyl alcohol, see p. 299) which is obtained in the destructive distillation of wood. This mixed spirit is equal to pure alcohol as a solvent, and it cannot be used for drinking purposes on account of the burning taste of the wood spirit. Moreover, the latter cannot be separated by any economical process. Methylated spirit is a treasure to the chemist.

Other Alcohols.—The term alcohol is applied in chemistry

to all hydrates of hydrocarbon radicals (page 297). The other alcohols are for the most part of no great practical importance.

ETHER.—ETHYL OXIDE.



It has already been shown (page 297) that the term ether is applied to the oxides of hydrocarbon radicals. Common ether, sometimes called *sulphuric ether*, because sulphuric acid is used in its preparation, is the most important member of the class, and the only one that needs description here. It differs from alcohol just as potassium oxide differs from potassium hydrate:

KOH Potassium hydrate $\text{C}_2\text{H}_5\text{OH}$ Ethyl hydrate (Alcohol).

K_2O Potassium oxide $(\text{C}_2\text{H}_5)_2\text{O}$ Ethyl oxide (Ether).

Like the other ethers, it is obtained by removing the elements of water from alcohol. This is generally effected by sulphuric acid.

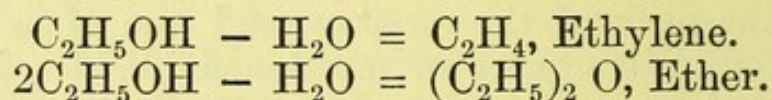
Preparation of Ether.—Experiment 1.—Put two fluid ounces of strong alcohol (or methylated spirit) into a good-sized flask, and add slowly one fluid ounce of strong sulphuric acid. Connect the flask with a receiver by means of a glass tube, close the opening remaining between the neck of the receiver and the glass tube by binding round it a piece of moistened bladder, in which some fine holes are pierced, and heat the flask *carefully* in a sand-bath till the contents of it assume a bubbling motion. Maintain the boiling of the liquid till about one ounce of the liquid is distilled over. In this experiment the liquor, as it is distilled, must be carefully cooled, because it is extremely volatile; it is therefore advisable, if possible, to surround the receiver with snow or ice. Care must also be taken not to bring any burning substance too near the vapours of the liquid which pass over, as they are *exceedingly inflammable*. The distilled, colourless liquid possesses a penetrating pleasant smell; it is called *crude ether*.

In order to purify it, shake it up in a bottle with half an ounce of water, and a little strong solution of potash; close

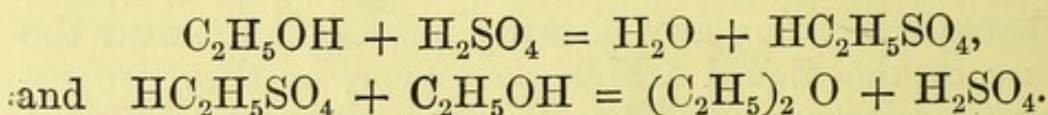
the bottle, and let it remain standing for an hour with the bottom upwards. Crude ether contains a mixture of water, alcohol, sulphuric acid and other impurities; these substances combine with the water and the potash added, and form with them the heavier liquid layer, which settles at the bottom of the phial. The very thin and mobile liquid floating above is *ether*, which separates, because it comports itself towards water in the same manner as oil does, and is dissolved by it only in very small quantity. If you now loosen the stopper of the inverted phial, the aqueous liquid will run out, while the ether remains behind. If the latter is required entirely pure it must be again distilled or rectified.

The most profitable way of preparing ether on a large scale is the following: Ten fluid ounces of sulphuric acid and twelve fluid ounces of alcohol are mixed together, and heated to the boiling point. While the mixture is still boiling, just so much alcohol is allowed gradually to drop in as there is ether distilled over. This is called the continuous process for the manufacture of ether.

It will be seen that the materials used for the preparation of ether are the same as for ethylene gas (page 184). But the proportion of sulphuric acid is smaller, and the dehydration of the alcohol is therefore not carried so far:



The curious fact that sulphuric acid will convert an indefinite quantity of alcohol into ether and water is explained by supposing that *hydrogen ethyl sulphate*, $\text{HC}_2\text{H}_5\text{SO}_4$, a salt of ethyl analogous to hydrogen-potassium sulphate, HKSO_4 , is first formed and then decomposed by another molecule of alcohol:



The water distils over with the ether, while the sulphuric acid remains to act on a fresh quantity of alcohol.

Experiment 2.—Pour some drops of ether upon the

hand; it will evaporate in a few moments, imparting to the hand a perceptible feeling of coldness. Ether is so *very volatile* that it boils at 95° F. (35° C.); therefore it must always be kept in tightly-closed bottles, and in a cool place.

Experiment 3.—Dip one piece of wood into ether, another into alcohol, and hold both to the flame of a candle; the ether *burns* with far greater briskness, and also with a much more luminous and a somewhat smoky flame. Its stronger illuminating power is simply explained by its containing a larger amount of carbon. The process in burning is the same as with alcohol; the ether being also converted into carbonic anhydride and water.

Experiment 4.—If you pour some drops of ether into a tumbler, and after some seconds, when the ether is converted into vapour, apply to it a burning taper, a sudden ignition ensues, accompanied by a small explosion. The vapour of ether forms, like hydrogen or marsh gas, when mixed with atmospheric air, a kind of *explosive gas*, and violent explosions have been occasioned by carrying lighted candles or lamps into those places where, owing to the breaking of a bottle filled with ether, its vapour has become diffused in the air.

Put a piece of *tallow*, or a few drops of olive oil, into a test-tube with some ether; both will entirely dissolve in it. But they are not soluble in alcohol or water. Therefore either may be advantageously employed for dissolving and separating such substances as will dissolve in it, but not in other liquids. Besides the fixed oils and fats, ether dissolves many resins, caoutchouc (india-rubber), gutta-percha, &c.

Ether is sparingly soluble in water, but may be mixed with alcohol in any proportion. It is very light (Sp. Gr. at 0° C. 0.736).

Other Ethers.—Ethers corresponding to many of the other alcohols are known, but few have any importance.

SALTS OF ETHYL.

It has been shown (page 300) that salts, sometimes called compound ethers, of the alcohol radicals can be prepared

which are in all respects comparable to metallic salts. Those of ethyl, which is the most important of alcohol radicals, are very numerous and interesting, but, as they do not possess much practical importance, we may pass them by in the simple practical study which we have in hand.

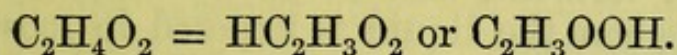
CHAPTER VI.

ORGANIC ACIDS.

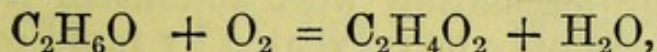
THE constitution of the simpler organic acids and their relation to the alcohols have already been explained (page 300). The student will find it well to read that section again before commencing the practical work of this chapter.

MONOBASIC ACIDS.

ACETIC ACID.



This acid, which is contained in vinegar, is by far the most important member of the series called the "fatty acids" (page 301). It is often obtained by the oxidation of alcohol, two atoms of hydrogen being replaced by one atom of oxygen.



but it is also formed during the destructive distillation of wood, and by other means.

Acetous fermentation.—*Vinegar.*—When weak wine or beer is exposed to the air, a peculiar organised ferment (page 341) soon appears in it, which has the remarkable power of effecting the oxidation of the alcohol into acetic acid. These "vinegar cells" differ from yeast cells in that they cannot develop except in the presence of oxygen, but that they are organised cells, capable of transplantation and increase, there is no doubt. Like yeast cells, they can only develop in presence of their appropriate food, and at a suitable temperature.

Experiment 1.—Mix in a glass vessel half an ounce of brandy with three ounces of water, and put in the liquid a slice of leavened bread, which has been previously

soaked in strong vinegar, or instead of it a little leaven (page 341); cover the vessel with a piece of perforated pasteboard, and put it in a place where the temperature is between 86° F. (30° C.) and 104° F. (40° C.); the spirituous liquor will, after some weeks, be converted into vinegar. This conversion does not take place in a closed vessel, as the oxygen of the air is indispensable to the process. Neither is any vinegar formed if you do not add the bread or the leaven. As a solution of sugar does not of itself pass into alcohol, neither does the alcohol of itself pass into vinegar.

In the same manner as with pure diluted alcohol, all other *alcoholic liquids*, as beer, wine, cider, &c., may, by receiving oxygen, be converted into vinegar, and it is well known that vinegar is frequently prepared from them. If, as is ordinarily the case, they contain gluten or lees in solution, then these substances afford food for the vinegar ferment, and the acidification ensues spontaneously, when the liquid is exposed in loosely covered vessels to a temperature of from 86° F. (30° C.) to 104° (40° C.). This acidification most readily occurs immediately after a spirituous fermentation, which has taken place at too high a temperature; for this reason, in the hot months of summer, the brewers and distillers find difficulty in keeping their fermenting wort and mash from turning sour, which can only be prevented by rapid cooling.

Experiment 2.—Fill two tumblers loosely with the stalks of grapes, and fill one entirely and the other only half full with wine, beer, or a mixture consisting of one part of brandy, one part of beer, and six parts of water. Put both vessels in a warm place, and once or twice every day pour the mixture from one vessel into the other, so that each may be alternately full and only half full of the liquid. The alcohol contained in the brandy will, in this manner, be much more rapidly oxidised into vinegar, because the liquid adhering to the grape-stalk is, in this state of fine division, surrounded by air, and thus has a far better opportunity of attracting oxygen from the latter. The effervescence taking place at the commencement is owing to the sugar contained in the beer and the grape-stalks, and which was first converted into alcohol

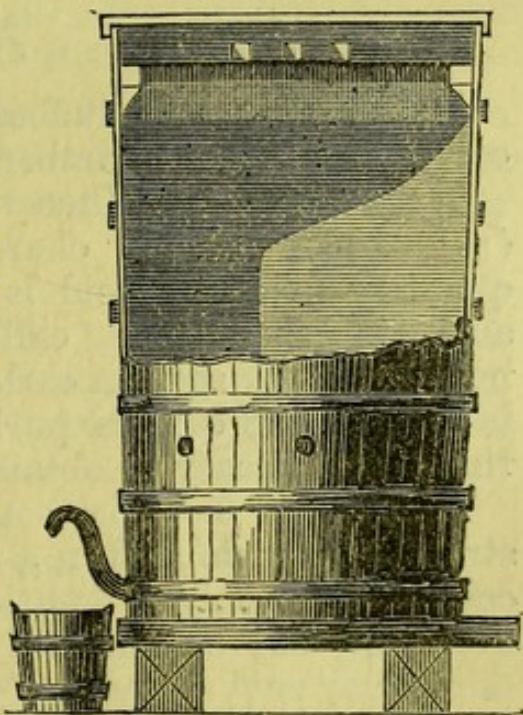
and carbonic anhydride. The alcohol thus formed is likewise afterwards changed into vinegar, and this is the reason why the vinegar thus produced is more acid, that is, richer in acetic acid, than that obtained by the former experiment.

Quick Method of making Vinegar.—The transition of alcohol into acetic acid takes place still more rapidly by subdividing the alcohol still further, or by exposing a still greater surface of the liquid to the air than in the way just described. This is effected in the following manner :

A tub four or five yards high is filled with shavings of beechwood, and is furnished with a perforated shelf, which is placed somewhat below the upper opening. Through each of the small holes a straw or a piece of packthread is passed, prevented from falling through by a knot at the upper end.

By this means an extreme division of the alcohol is effected, as when it is poured in at the top, it only trickles slowly down through the holes by means of the straw or packthread, and then diffuses itself over the shavings, forming a very thin liquid layer, which presents to the air a surface many thousand times more extensive than was produced by any former method. Several large holes are bored round the lower part of the tub, and likewise in the perforated shelf; glass tubes are fitted

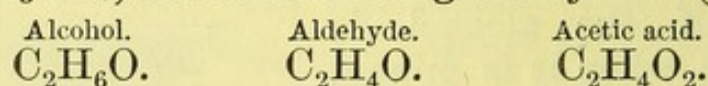
Fig. 104.



into the holes made in the latter, in such a manner that the liquid, when poured into the top, may not run off through them. A free circulation of air is hereby produced, the cooler air enters by the openings in the tub, gives up its oxygen to the alcohol diffused over the shavings, and, in consequence of this oxidation,

or slow combustion, so much heat is evolved in the interior of the tub that the temperature rises to 104° F. (40° C.). The air hereby becoming warmer, and consequently lighter, passes out of the tub through the glass tubes in the shelf, from an eighth to a fourth poorer in oxygen than when it entered. Strong vinegar is used to assist the fermentation in this process, the tub and shavings having previously been moistened with it, and a certain quantity being also added to the mixture of spirit which is to be converted into vinegar. In such a tub (*vinegar generator*), warmed spirit, beer, wine, &c., may be converted into vinegar in a few hours, by being passed through the cask three or four times; hence this is called *the quick method of making vinegar*.

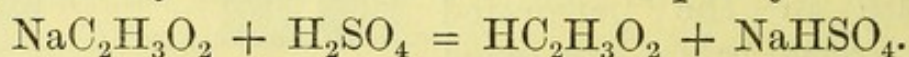
Aldehyde.—During this oxidation of alcohol, an intermediate compound called aldehyde is first formed, and then oxidised to acetic acid. It is regarded as *hydride of acetyl*, (C_2H_3OH), acetic acid being the *hydrate* (page 300):



Its peculiar and suffocating odour can often be perceived in vinegar chambers.

Wood Vinegar.—When wood is distilled in close vessels (page 188), besides charcoal, tar and gas, a certain quantity of acetic acid is always formed. This impure acid is sometimes called *pyroligneous acid*. When neutralised by sodium carbonate, *sodium acetate*, $NaC_2H_3O_2$, is formed, and can be purified by crystallization. From this salt it is easy to obtain pure acetic acid.

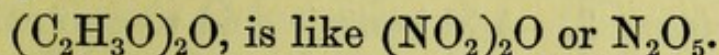
Experiment 3.—Mix some dry sodium acetate with strong sulphuric acid in a flask connected with a cool receiver, and apply a gentle heat. Acetic acid will distil over, and may be collected in a state of purity:



Pure acetic acid is an ice-like solid below 63° F. (17° C.), and is therefore sometimes called *glacial acetic acid*. It boils at 244° F. (118° C.). The strongest English vinegar contains about 5 per cent. of the true acid.

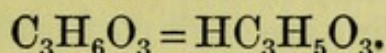
Acetic anhydride, or *Acetyl oxide*, $(C_2H_3O)_2O$, can only be

prepared with difficulty. In constitution it is analogous to nitric anhydride :

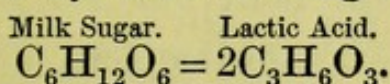


Acetates.—Those of lead (sugar of lead) and sodium are among the most important. Acetates can be recognised by warming them with sulphuric acid, when the odour of vinegar will be apparent.

LACTIC ACID.



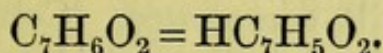
Lactic fermentation.—When milk is kept for some hours in a warm place, the milk-sugar undergoes a fermentation analogous to those already described. It is converted entirely into lactic acid, which renders the milk sour, and precipitates the casein as *curd* (page 328). The nature of the change is shown by the following formula :



Lactic acid can only be purified with difficulty. It is a colourless, oily liquid, which is decomposed by heat. A variety of lactic acid known as *sarcophilic acid* is found in the juice of flesh.

Butyric fermentation.—If the fermentation of milk is pushed past the stage in which lactic acid is formed, *butyric acid* is produced by a distinct organised ferment. The odour and taste of rancid butter and putrid cheese are partly due to this acid. It is one of the fatty acids.

BENZOIC ACID.



Experiment 1.—Put some powdered gum benzoin into a saucer, cover it with a sheet of white filtering-paper on which rests a cone of writing-paper. Apply a very gentle heat to the saucer by means of a sand bath, and after an hour or so the inside of the cone will be found to be covered with beautiful white crystals of benzoic acid. Good gum contains about 14 per cent. of the acid, and the acid being volatile passes in vapour through the porous

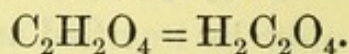
paper, and condenses on the cone. The filter paper absorbs some volatile oil which is given off with the acid.

Experiment 2.—Add benzoic acid to a solution of sodium carbonate until it no longer dissolves. Effervescence will take place from the escape of CO_2 , and *sodium benzoate*, $\text{Na C}_7\text{H}_5\text{O}_2$, will remain in solution, and may be obtained in the solid state by crystallization. From this salt other benzoates may be prepared by double decomposition.

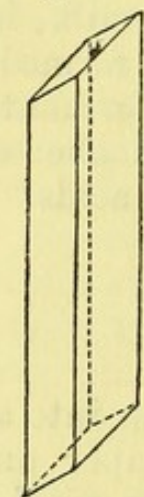
Experiment 3.—Add solution of ferric chloride (rendered neutral by ammonia) to solution of sodium benzoate. A buff-coloured precipitate of *ferric benzoate* will be produced.

Experiment 4.—To a tolerably strong solution of sodium benzoate add hydrochloric acid. Pearly crystals of benzoic acid will be thrown down. Benzoates can be decomposed in this manner by very many acids. The precipitation of the acid is due to its sparing solubility. It requires 200 times its weight of cold water for solution.

OXALIC ACID.



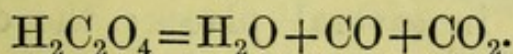
This important dibasic acid corresponds to an oxide of carbon, C_2O_3 , intermediate between carbonic oxide and carbonic anhydride. This oxide is however unknown. The relation of oxalic acid to ethylene alcohol has already been shown (page 301).



Experiment 1.—Heat with free access of air, in a porcelain dish, one-fourth of an ounce of sugar, mixed with one and a half ounce of concentrated nitric acid, and one ounce of water. In a short time a strong evolution of yellowish-red fumes (N_2O_3) will commence. Continue boiling until these vapours cease, and then put the liquid in a cool place; colourless crystals (oblique rhombic prisms) of oxalic acid will separate, which must be purified by recrystallization. They have a strong acid reaction, and are poisonous. Starch, or even saw-dust (impure cellulin), may be substituted for sugar in this experiment. The crystals contain $\text{H}_2\text{C}_2\text{O}_4, 2\text{H}_2\text{O}$.

Experiment 2.—Put about half an inch of crystallized oxalic acid into a test tube. Moisten with strong

sulphuric acid, and carefully heat the mixture; a gas will be evolved. Let this pass through solution of potash contained in another test-tube. One half of the escaping gas is absorbed by the potash; this is carbonic anhydride (CO_2). The other half escapes through the open tube, and burns, when kindled, with a bluish flame; this is carbonic oxide (CO). The carbonic oxide will even burn at the mouth of the tube if the carbonic anhydride remains with it. When the evolution of gas ceases, there will be found in the first test-tube common sulphuric acid; consequently, the sulphuric acid has removed water from the oxalic acid. The oxalic acid, when it loses its water, is resolved into the two gases just mentioned, in the following manner:



Experiment 3.—Place some crystals of oxalic acid upon a piece of platinum foil, and hold them in the flame of a spirit-lamp. They melt, take fire, and burn *without becoming black* or leaving any residue. The product of the combustion is carbonic acid; $\text{H}_2\text{C}_2\text{O}_4$ and O (from the air) are converted into $2\text{CO}_2 + \text{H}_2\text{O}$.

Experiment 4.—Neutralise a strong solution of oxalic acid with a strong solution of potassium carbonate or hydrate; neutral potassium oxalate ($\text{K}_2\text{C}_2\text{O}_4$), an easily soluble salt, is formed. If you now add as much more oxalic acid, hard crystals will be deposited on cooling, which have an acid reaction; they are called acid oxalate of potassium, or *hydrogen potassium oxalate*, HKC_2O_4 . The acid oxalate is likewise formed by the vital process in many plants, and it is found abundantly in the leaves of the wood-sorrel (*Oxalis*), from which it may be obtained. The acid salt is far less soluble than the neutral.

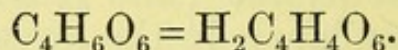
Experiment 5.—Heat some potassium oxalate upon platinum foil; it will be converted into potassium carbonate. The oxalate loses carbonic oxide: $\text{K}_2\text{C}_2\text{O}_4 = \text{K}_2\text{CO}_3 + \text{CO}$. Many other oxalates are converted into carbonates when treated in this way.

Experiment 6.—Agitate a little calcium sulphate with water and let the liquid settle; the decanted water contains a small quantity (about $\frac{1}{400}$) of the sulphate in

solution. If a solution of oxalic acid is poured upon this solution, you will soon obtain a precipitate of calcium oxalate. The decomposition takes place more rapidly and perfectly when the oxalic acid has been previously neutralised by ammonia. Oxalic acid, or ammonium oxalate, is the best test for calcium salts.

Experiment 7.—Pour a solution of ferrous sulphate on a piece of white blotting-paper; when this has imbibed the liquid, spread it over some ammonia. The ammonia precipitates green ferrous hydrate on the paper, which soon absorbs oxygen, and is converted to the brown ferric hydrate. In a similar manner, cotton, and other fabrics, are often dyed brown or yellow. When it is dry, mix some oxalic acid with water into a thin paste, and dot the paper with it in several places; the colour will soon disappear from these spots, and you obtain a white pattern on a yellow ground. Oxalic acid dissolves ferric oxide, and both are removed by washing. Upon this is founded an important use of this acid in calico-printing, as likewise its application for the removal of ink-spots from linen or paper. One of the principal constituents of ink is an iron salt, which being dissolved by the oxalic acid, the black colour of the ink disappears also. This explains why, on removing ink-spots by oxalic acid or salt of sorrel (which acts in the same manner as the free acid) from yellow and brown articles of dress, the colour of which often depends upon iron, the colour of the stuff also disappears.

TARTARIC ACID.



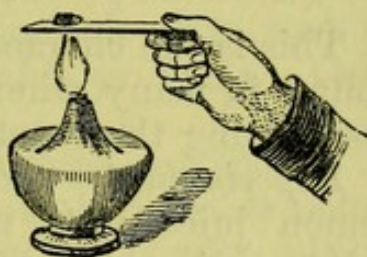
Tartaric acid is dibasic. It is generally prepared from tartar or argol, which is obtained in large quantities from the wine countries, where it is deposited from wines in the casks, as a white or reddish crust of impure hydrogen potassium tartrate. Tartaric acid might be very easily obtained from this salt by means of sulphuric acid; but then two soluble substances would be obtained, which could not well be separated from each other. Therefore the tartar is first converted into calcium tartrate, and this when treated with sulphuric acid yields tartaric acid,

which can be purified by crystallization. The calcium remains as the nearly insoluble sulphate.

Tartaric acid has very much the appearance of a salt; it crystallizes in colourless oblique prisms, which are permanent in the air, and have a very acid taste.

Experiment 1.—Place a small crystal of tartaric acid upon a piece of platinum foil, and heat it over the flame of a lamp; it will first melt, then become brown, and finally black, and emit at the same time a smell somewhat like that of burnt sugar. If, during the process of charring, you hold over the acid a dry, cold glass vessel, it will become lined with globules of water; consequently the acid contains oxygen and hydrogen. The dark residue resembles charcoal, but it is more certainly recognised as such by its burning completely at a higher heat. Accordingly tartaric acid, when heated, behaves like sugar.

Fig. 106.



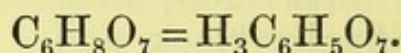
Experiment 2.—Pour a little warm water over some tartaric acid; it will dissolve therein, for it is *readily soluble* in water. If you dilute the solution with more water, and put it aside in a moderately warm place, slimy flakes will be deposited, and the acid taste will gradually be lost—it decomposes. Many other organic compounds, when they are diluted with water, decompose after a time.

Experiment 3.—Neutralise a solution of potassium carbonate with a solution of tartaric acid; carbonic anhydride escapes; the liquid, however, remains clear, because the neutral potassium tartrate ($K_2C_4H_4O_6$) formed is an easily soluble salt. But by adding yet more tartaric acid, the liquid becomes turbid, especially on stirring, and deposits a quantity of small transparent crystals, which are difficultly soluble in water, have an acid taste, and contain less potassium than the neutral salt. These crystals are called acid tartrate of potassium, or hydrogen potassium tartrate, $HKC_4H_4O_6$; commonly, tartar, or when pulverised, cream of tartar. The salts of potassium may accordingly be used as *tests for tartaric acid*.

Experiment 4.—If you heat the crystalline powder of tartar, obtained in the last experiment, on platinum foil,

it will, like the tartaric acid, become black, and burns with an empyreumatic odour; but there will, however, finally remain a white powder, which has an alkaline taste, a basic reaction, and which, on the addition of an acid, will effervesce. It is potassium carbonate.

CITRIC ACID.



This is the characteristic acid of lemons, but it is also found in many other fruits. It is tribasic, as is shown by its forming three potassium salts.

Experiment 1.—Add chalk in small portions to some lemon juice, until no further effervescence occurs. The white sediment consists of impure *calcium citrate*. It may be carefully decomposed by sulphuric acid (avoiding excess), the free citric acid filtered from the insoluble calcium sulphate, and crystallized. It is however somewhat difficult to obtain it pure by this process.

Experiment 2.—Citric acid resembles tartaric acid in its appearance and many of its properties. The effect of *heat* on the crystals may be tried.

Experiment 3.—Add some lime water to a solution of citric acid. No precipitate, or only a very slight one, is formed. Boil the liquid and an abundant precipitate will appear, but will again disappear as the liquid cools. *Calcium citrate is more soluble in cold water than in hot; a remarkable exception to the general rule.*

GALLIC ACID.—TANNIN.

Gallic Acid, $\text{C}_7\text{H}_6\text{O}_5$. Tannin, $\text{C}_{14}\text{H}_{10}\text{O}_9$.

Closely allied monobasic acids.

Tannin, Tannic Acid, Gallotannic Acid.

Experiment 1.—Take a large test-tube with a small hole in its end, plug the hole with cotton wool, and fill the tube half full of the best Aleppo *gall nuts* in fine powder. The gall nuts found on English oaks are very inferior. Support the tube upright, and pour in some common ether. Commercial ether always contains some water and alcohol, both of which in small quantity are necessary for

the success of the operation. The ether soon percolates through the gall nuts, and may be received in a small bottle. Repeat the addition of ether several times, and then cork up the bottle and allow it to remain at rest for some time. The liquid will separate into two layers, the upper one of which consists mainly of ether, which may be drawn off with a pipette,* and the ether purified by distillation. The lower layer consists chiefly of an aqueous solution of *tannin*, or *tannic acid*, a peculiar astringent substance which is contained in gall nuts, and in less quantity in a great many vegetable products. Oak-bark contains a good deal of it.

The aqueous solution may be evaporated to dryness at a gentle heat, when the tannin is obtained as an amorphous mass of strong astringent taste. It is very soluble in water and alcohol, but insoluble in pure ether.

The chief importance of tannin arises from its power of combining with gelatin and gelatigenous substances, and forming with them a substance, *leather*, which does not putrefy. The process of *tanning* may be illustrated in the following manner:

Experiment 2.—Add a solution of gelatin to a dilute solution of tannin, or to a decoction of oak bark or gall nuts; a white flocculent precipitate will be formed.

Experiment 3.—Take a clean, dry piece of skin, weigh it, and suspend it for some days in a solution of tannin, or a decoction of bark. It will gradually remove the whole of the tannin, and be converted to an imperfect leather. When dried and weighed it will be found to have gained in weight, the increase affording a rough measure of the amount of tannin that the solution contained.

For the tanning of thick hides many months are often required.

Experiment 4.—A solution of tannin gives with *ferric* salt a blue-black precipitate, which is the colouring matter of ordinary *ink*. Pure *ferrous* salts give a white precipitate with tannin, which absorbs oxygen from the air and turns black. The reason why so many inks become

* Pipettes are tubes drawn out to a fine jet at one end. They are very useful for sucking up small quantities of liquid. They are often made with a bulb in the middle.

darker when the writing is exposed to air is that a ferrous salt has been used in their manufacture.

Experiment 5.—Writing Ink.—Take one pound of bruised Aleppo gall nuts, six ounces of green vitriol (ferrous sulphate), six ounces of gum arabic, a few drops of carbolic acid (or creosote), and one gallon of cold water. Mix them in a stone jar, and stir at intervals for three weeks; then let the ink settle and decant it for use. The gum gives thickness to the ink, the carbolic acid keeps it from turning mouldy.

Experiment 6.—Tea contains a good deal of tannin. Experiments 2, 3, and 4 may be repeated with a decoction of tea. Tea made with waters which contain iron has an inky appearance.

Gallic Acid.

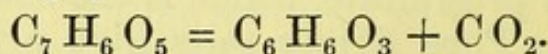
Experiment 7.—To prepare gallic acid, moisten some powdered gall nuts with water, and keep them for a month in a warm place (about 20° C., or 68° F.), replacing the water as it evaporates, and removing the mould which forms. Press the mass in a calico bag, and afterwards boil it with water and filter while hot. As the solution cools, silky crystals of gallic acid will deposit.

Gallic acid is distinguished from tannin by its power of crystallizing and by the fact that it does not precipitate gelatin. It is useless in tanning.

Experiment 8.—Gallic acid gives with ferric salts an even more intense blue-black colour than tannin does. Its solution affords a most delicate test for those compounds.

Pyrogallic Acid, or Pyrogallol, C₆ H₆ O₃.

Experiment 9.—Heat some gallic acid, very gently, in the apparatus used for preparing benzoic acid (page 353). After some time beautiful silky crystals of pyrogallic acid are found in the paper cone:



Pyrogallic acid is used in photography. Its alkaline solution absorbs oxygen rapidly and turns brown; it is therefore much used in gas analysis. In constitution it is allied to phenol.

CHAPTER VII.

FATS AND FIXED OILS, SOAP, &c.

GROUP i.—Non-drying Fats and Oils (Glycerides).

GROUP ii.—Drying Oils (also Glycerides).

GROUP iii.—Wax, Spermaceti, &c. (not Glycerides).

GROUP i.—Non-drying Fats and Oils.

THROUGHOUT the vegetable and animal kingdoms a class of oily substances is distributed which is of very great interest and importance. The members of this class are generally called *fats* when solid, and *oils* when liquid at ordinary temperatures. They are almost invariably mixtures in various proportions of two or more definite compounds. They possess in common certain well-marked peculiarities, which may be shortly enumerated.

1. They are all *compounds of carbon, hydrogen and oxygen only*.

2. They are *lighter than water*.

3. They are *non-volatile*, being destroyed and converted into gases when heated above a certain point.

4. They are *insoluble* in water, *scarcely soluble* in alcohol, *very soluble* in ether, benzene and carbon disulphide.

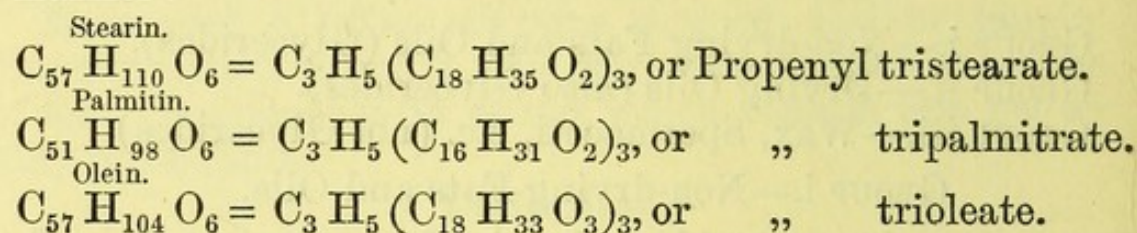
5. They are *fusible* at a temperature much below that of boiling water.

6. When *boiled with caustic alkali* they all yield glycerine, and peculiar salts of the alkaline metal (*soaps*: the change is called *saponification*).

Proximate constituents of the members of this group.—The most important of these are: *stearin*, $C_{57}H_{110}O_6$, a white solid, the chief constituent of suet, tallow, and lard, contained also in some vegetable oils; *palmitin*, $C_{51}H_{98}O_6$, solid, not so hard as stearin, found in palm oil, and in many animal and vegetable fats; and *olein*, $C_{57}H_{104}O_6$, liquid, the

chief constituent of olive and many other oils, and also present in most fats.

These three substances are called *glycerides*, because they are salts of the triad radical *glyceryl* or *propenyl*, C_3H_5 (page 294), of which glycerin is the hydrate, $C_3H_5(OH)_3$. Each of them contains the monad radical of a distinct acid; stearin, of *stearic acid*, $HC_{18}H_{35}O_2$; palmitin, of *palmitic acid*, $HC_{16}H_{31}O_2$; and olein, of *oleic acid*, $HC_{18}H_{33}O_2$. Their constitution may be represented in the following manner:



Palmitic and stearic acids are members of the formic or "fatty acid" series; oleic acid, of the acrylic series (page 302).

ANIMAL FATS AND OILS.

Experiment 1.—Boil some fat pork, cut up into small pieces, for some time in a little water, and while the soft mass is still hot, press it in a linen cloth; a fat oil will float on the surface, but it is fluid only at a temperature of about $86^\circ F.$ ($30^\circ C.$); below this temperature it congeals into a solid, yet soft, white substance. This is lubricating to the touch, and produces greasy spots on paper. Those kinds of fat, which, at the common temperature, have a soft unctuous consistency, are called *lard*; and the cellular membrane and skin remaining in the cloth, and saturated with fat, are called *scraps*.

The suet of mutton, when treated in the same way, yields a fat which, when hot, is also fluid, like oil, but which, when cooled only to about $97^\circ F.$ ($36^\circ C.$), congeals, and then forms *tallow*, a harder substance than lard. By boiling and roasting, we can melt out fat from many animal tissues, especially from those of the domestic animals, in which we are able to produce a great quantity of fat by keeping them confined, and giving them a plentiful supply of food. The fats obtained by boiling

with water are white, as thereby they do not become heated above 242° F. (100° C.); while those obtained by roasting have a yellow or brown colour (brown butter, gravy of roast meat, &c.), because in this case a portion of the fat becomes burnt by being subjected to a stronger heat—to a heat, perhaps, even above 572° F. (300° C.).

These substances consist mainly of stearin, palmitin and olein, the first predominating. When lard is strongly pressed, olein is obtained from it (lard oil).

Train, or whale oil, is obtained from the blubber of the whale. *Sperm oil*, from cavities in the skull of the sperm whale.

Cod liver oil, obtained from the liver of the cod, contains minute quantities of bromine, iodine, sulphur, phosphorus, and free acids, in addition to stearin, palmitin, and olein. To the bromine and iodine the value of the oil in medicine is often ascribed.

Cream.—Butter.—That singularly complex and important liquid *milk* owes its opacity to myriads of minute globules of oil which float in it. Under the microscope these are seen to be perfectly transparent, for milk is opaque, just as powdered or ground glass is opaque, only because the rays of light which fall on it are scattered by the irregularity of the globules or fragments. When milk is kept for a time, most of the oil globules rise to the surface and constitute *cream*. They do not run together, because each is enclosed in a delicate membrane; but when the cream is violently agitated (*churned*), the membranes are broken and a solid mass—*butter*—is obtained. *Skim-milk* is milk from which the greater part of the cream has been skimmed. *Butter-milk* is the liquid residue left after butter has been made from cream. Both consist chiefly of *milk-sugar* and *casein* (page 328) and mineral salts, with much water. In the modern *cream separator*, the milk, rotating in drums with great velocity, is divided in a few minutes into *cream*, which goes to the outer edge and is skimmed off by a sharp-edged pipe, and *skim-milk* which is separately received.

Fresh butter consists chiefly of olein, palmitin and stearin, with small quantities of other glycerides, and of albuminoid bodies (casein, &c.). When kept, the albu-

minoids begin to putrefy, and a kind of fermentation takes place, which produces soluble fatty acids. To some of these fatty acids (butyric, caprylic, capric, &c.) the disagreeable taste and smell of rancid butter is due. The change can be to a great extent prevented by *salt*, which acts as an antiseptic, or by purifying the butter by melting it in hot water, and removing the solid mass when cold. The latter process, however, injures the delicate flavour of the butter.

VEGETABLE FATS AND OILS.

Almond oil.—*Experiment 1.*—Break open an almond, and squeeze the white meat together by means of the finger-nail; small drops of fluid will be expressed, which are slippery to the touch, and render blotting-paper greasy and transparent. This liquid is called *sweet oil of almonds*. If the almonds are first pounded, and then subjected in a cloth to strong pressure, we shall obtain more than one-fourth of their weight of oil of almonds. A great many plants contain a similar oily juice, especially in their seeds, and from many of the latter, oils are obtained by pounding and expressing. They occur in small quantity in almost all plants, even in those where we should not expect to find any; for instance, in different kinds of corn, grasses, &c.

Palm oil is obtained from the fruit of certain kinds of palm. It is yellow and has the consistence of butter. It consists chiefly of palmitin, with olein, &c. It is largely used in the manufacture of soap and candles, and also, with tallow and a little soda, for greasing the wheels of railway carriages. For the two former purposes it is bleached by exposure to air in a melted state for ten or fifteen hours.

Cocoa-nut oil is a soft white solid, somewhat like palm oil.

Olive oil, expressed from the fruit of the olive tree, is too well known to require description. A simple experiment shows that it is a mixture.

Experiment 2.—Expose some olive oil to the cold of a winter's night, or surround it for some time with ice. It will become semi-solid, and will be full of granules.

These granules consist of *palmitin*. Subject the mass to gentle pressure in calico without allowing the temperature to rise; a clear oil will be expressed, which does not solidify at 32° F. (0° C.). This consists almost entirely of *olein*.

Colza oil and *Rape oil* are obtained from the seeds of several species of *brassica*, to which genus the cabbage, turnip, and rape belong. They are said to contain the glyceride of a peculiar acid, *brassic acid*.

The following experiment illustrates the mode in which rape oil is *refined* or purified.

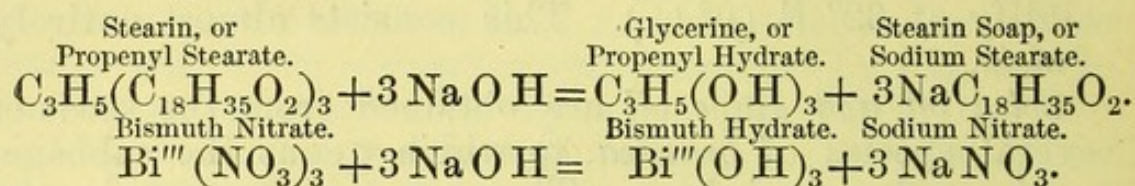
Experiment 3.—Mix one ounce of crude rape oil with eight drops of common sulphuric acid, and shake it frequently; in half an hour add half an ounce of water; again shake the mixture briskly, and set it aside for some days, when the oil floating on the surface will be freed from mucilage (purified). The slimy parts, charred by the sulphuric acid, and rendered insoluble, are found settled in the water at the bottom of the vessel. The sulphuric acid still adhering to the oil is removed by repeated washing with water. Sulphuric acid chars, as is known, all organic substances, some (for instance, mucilage) easily, others (for instance, oil) with difficulty; if just enough sulphuric acid, therefore, is added to the oil to char the mucilage, then the mucilage only is destroyed, and the oil remains undecomposed. A larger quantity of sulphuric acid would also attack the oil.

Castor oil, from the seeds of *Ricinus communis*, consists chiefly of the glyceride of a peculiar acid, *ricinoleic acid*. It also contains small quantities of a peculiar alkaloid, *ricinine*, discovered by Tuson. The oil is remarkable for being easily soluble in alcohol.

SAPONIFICATION.—SOAP.

When any one of the preceding fats and oils—in fact, any glyceride, or mixture of glycerides—is exposed at a suitable temperature, different with different oils, to the action of a strong solution of caustic soda or potash (alkaline lye), a double decomposition takes place, glycerine is formed, and with it a sodium or potassium salt of the acid, or acids of the oil. These salts are called *soaps*,

and the change is called *saponification*. The following formulæ will show how similar the action is to some of the changes of mineral chemistry :



Soaps are in fact salts of the higher fatty acids. The ordinary *hard soaps* are *sodium* salts. *Soft soaps* are *potassium* salts. These soaps are soluble in water, alcohol, &c., but others, such as *lead soap* (lead plaster), *calcium soap*, &c., can be prepared which are insoluble in water. It will readily be understood that ordinary washing soaps are mixtures of stearate, palmitate and oleate of sodium.

Hard soap.—*Experiment 1.*—Make first a strong lye with one drachm of caustic soda and one ounce of water, and next, a weak lye, with one drachm of caustic soda and two ounces of water. Boil the latter gently with an ounce and a half of beef-suet, for half an hour, in a vessel only half filled with the mixture, and then add the strong lye gradually while the boiling continues. The fat and lye gradually unite and form a uniform mass, of a gluey consistency, which after a time becomes thick and frothy. If a drop of this, when pressed between the fingers, presents firm white flakes, then add half an ounce of common salt, boil for some minutes, and let the whole mass cool quietly. We obtain a firm mass (*soap*) and a watery liquid, in which the common salt and some free soda remain dissolved. If the soap, when boiled with distilled water, forms a turbid solution, it still contains some unsaponified fat, in which case add to it some weak lye, and continue boiling until the sample gives a clear solution in water; add again some common salt, and let it cool. The soap prepared in this manner has the same composition as common *house-soap*. Palm oil or cocoa-nut oil is often used partly or entirely to supply the place of tallow; the palm oil because it is cheaper than tallow, and the cocoa-nut oil because it communicates to the soap the property of forming a very strong lather.

Experiment 2.—Repeat the former experiment, using

olive oil instead of tallow ; hard soap is likewise obtained (*oil- or Marseilles soap*).

Soft soap.—*Experiment 3.*—Prepare again some oil-soap, as above described, but instead of soda, use potash-lye, which is prepared from caustic potash and water, and omit the addition of common salt ; the glutinous mass does not then pass by boiling into a hard soap, but, after sufficient evaporation of the water, yields a soft mass. This kind of soap is frequently employed in print-works for the cleansing of coloured fabrics. If train oil, hemp-seed oil, or linseed oil is used instead of olive oil, a darker-coloured soft soap is obtained.

Ammonia acts far more feebly than potash and soda upon fats. If some oils are shaken up with ammonia, thick white mixtures, called *liniments*, are obtained, which are often applied by friction to the skin.

Experiment 4.—The action of common salt in soap-making may be seen by trying to dissolve hard soap in salt water ; no solution takes place, not even on boiling, *for soap is insoluble in salt water*, and likewise in strong lye ; therefore soap may be precipitated from a solution in water by the addition of common salt. This method of separation is sometimes employed on the large scale, since it yields a purer soap than when the water is removed by evaporation ; for, in the latter case, glycerine, surplus of lye, and the impurities contained in the lye or fat, remain mixed with the soap, while by the former method they are retained in the liquid.

Experiment 5.—Dissolve some hard soap in hot soft water, and add white vinegar, a drop at a time, till the liquid is thoroughly turbid. The acetic acid liberates the fatty acids of the soap, which after a time rise to the surface of the liquid and float there, being insoluble in water. The liquid now contains *sodium acetate*.

Washing with soaps.—Soaps have two important properties :—1st, they dissolve fat and oils ; 2nd, they are very easily resolved, merely by mixing with much water, into an acid salt and free alkali ; the latter dissolves, as is well known, many organic substances, while the former effects by its lubricity an easy washing away of the dissolved matter from other substances. On these two pro-

perties depends the application of soap in washing. The separated acid salt of the fatty acids both diminishes the action of the free alkali, and keeps the articles pliant which are washed with soap, while they would become rigid if they were cleansed with caustic alkalies alone. To prevent the shrinking of woollen articles, wash them with a weak solution of sodium carbonate instead of with soap.

Soap and Alcohol.—*Experiment 6.*—Pour one ounce of alcohol upon one ounce of the shavings of soap; the soap is completely dissolved on heating in the water-bath, but the solution congeals on cooling to a transparent *jelly*. This jelly-like soap, mixed with camphor and ammonia, is called *opodeldoc*. The white stars separating from this consist of crystallized sodium stearate. All soaps prepared from solid fats (rich in stearin) behave like tallow soap.

Experiment 7.—Dissolve one drachm of Naples soap in half an ounce of alcohol; this solution does not coagulate on cooling; it forms a *tincture of soap*. By evaporation, a diaphanous soap is obtained (*transparent soap*). All the soaps made from the fluid fats (rich in olein) act like the Naples soap.

Experiment 8.—If some lime-water is added to a solution of soap in water, a precipitate of insoluble calcium soap is formed; this explains why spring-water, which generally contains calcium and magnesium salts (hard water), neither dissolves soap nor lathers with it, and accordingly cannot economically be used for washing. Hard water may, however, be softened by the addition of sodium carbonate (common washing soda), which precipitates the calcium and magnesium as insoluble carbonates. Of course the metals can also be removed by soap, if enough is used, but washer-women use soda instead because it is cheaper.

Yellow soap.—The tallow and palm oil used in soap-making is often mixed with common rosin, which at a high temperature combines with alkalies even more easily than the fats. The soap so obtained is called yellow soap. It is cheap, but has a disagreeable odour.

Marine soap is made from cocoa-nut oil. It is soluble in

somewhat dilute brine, and is therefore used for washing in sea-water.

Glycerine.—*Experiment 9.*—In a deep basin, heat together 9 parts of olive oil (or any other glyceride), 5 parts of finely-powdered litharge and a little water. A little more water must be added from time to time to replace the loss from evaporation. After a time the oil disappears, and a firm mass insoluble in water is obtained and a watery liquid. The mass is the insoluble lead soap generally known as lead plaster. It consists, if olive oil has been used, chiefly of *lead oleate*. The watery liquid is a solution of glycerine containing some dissolved lead. It must be separated from the liquid, and the latter thoroughly saturated with sulphuretted hydrogen by passing a stream of the gas through it. This throws down the lead as sulphide, and the filtered liquid, evaporated to a syrup on the water bath, yields *glycerine*. It generally retains, however, a small quantity of lead.

A much better process for the preparation of glycerine is generally adopted on the large scale. It depends on the fact that superheated steam decomposes glycerides into glycerine and the fatty acids, both of which are carried over with the steam and readily condense in separate layers. The process is usually applied to solid fat (tallow, palm oil, &c.), and the solid fatty acids, stearic and palmitic, are used for the manufacture of the so-called *stearin*, *palm*, or *composite candles*, now so largely used.

Properties of Glycerine.—Pure glycerine is a colourless syrupy liquid of sp. gr. 1.28. It mixes with water and alcohol in all proportions, but is insoluble in ether. It has a very sweet taste, whence its name (from $\gamma\lambda\upsilon\kappa\acute{o}\varsigma$, sweet). Distilled with steam, or at a reduced pressure, it passes over unchanged, but under ordinary circumstances the greater part is decomposed, and a substance of pungent and irritating odour, called *acrolein*, C_3H_4O , is produced.

Experiment 10.—Boil a little glycerine in a test-tube, the irritating odour of *acrolein* will soon be perceived. If some hydrogen potassium sulphate is mixed with the glycerine, acrolein is produced in much larger quantities,

but, unless elaborate precautions are taken, it is impossible to remain in the laboratory while the experiment is in progress, as the vapour, even in very minute quantity, irritates the eyes insupportably. Acrolein, treated with silver oxide, yields the silver salt of *acrylic acid*, $\text{AgC}_3\text{H}_3\text{O}_2$ (page 302), of which acid acrolein is the aldehyde.

When glycerine is heated with fatty acids under pressure, oils or fats identical with the proximate constituents of the natural fats and oils (page 361) are obtained, water being formed at the same time.

GROUP ii.—Drying or Varnish Oils.

We have seen that many seeds, when submitted to pressure, yield fixed oils analogous in their nature to the animal fats and oils. A certain number of these oils possess the property of *drying*, that is, of becoming solid and hard when exposed to the air. The change is attended with the absorption of oxygen, but is not well understood. It renders them very valuable in the manufacture of paints and varnishes. The oils of linseed, hemp-seed, and poppy-seed, and the oil of walnut kernels (nut oil), are among the most important of this class. Linseed oil, which is more largely used than the others, is the only one which need be studied here.

All the drying oils are mixtures of glycerides. Linseed oil contains the glyceride of a peculiar acid, *linoleic acid*. They are, of course, all capable of saponification.

Linseed oil.—The well-known linseed, the seed of common flax, yields, on being subjected to pressure, a yellow oil, equal to one-fourth of its own weight, which is gradually bleached by long exposure to the sunlight. It is most frequently used in oil varnishes.

Experiment 1.—Add to an ounce of linseed oil a quarter of a drachm of litharge, and half a drachm of acetate of lead; put the mixture in a warm place, and frequently shake it. The liquid, clarified by settling, now dries much quicker than it would have done before; it is used in the manufacture of the common *linseed oil varnish*, which, mixed with colours, is generally used for imparting a gloss to wood, metal, &c. The so-called oil-cloth is cotton cloth smeared with coloured linseed oil;

oil-silk is varnished silk. The oil is commonly prepared, on a large scale, by heating one hundred pounds of linseed oil with three pounds of litharge, and maintaining the mixture for an hour at a temperature of 212° F. (100° C.). A stronger heat renders the varnish darker and thicker, and, besides, might easily cause it to boil over and take fire. The slimy, dingy white sediment which remains after both processes is a combination of mucilaginous substances with oxide of lead. All oils contain, in the unpurified state, mucilaginous (gummy and albuminous) substances, which retard the drying; these are rendered insoluble by oxide of lead. Varnish is, accordingly, linseed oil free from mucilage.

Experiment 2.—The difference between the drying and non-drying oils may be shown by smearing two pennies, one with olive, the other with linseed oil, and allowing them to remain for several days in a warm place. The linseed oil will dry up into a thin transparent solid layer.

Experiment 3.—Heat some linseed oil over a lamp, and test the temperature of it occasionally by a thermometer. At first the heat rapidly rises to 212° F. (100° C.), and remains for some time at that temperature, during which time the oil boils moderately; this behaviour is occasioned by all crude oil containing watery particles, which evaporate at 212° F. (100° C.). As soon as these have volatilised, the temperature soon rises to 300° C. (572° F.), or even higher, when the oil begins to boil for the second time, but emits now a white smoke having a very disagreeable odour. This vapour consists principally of illuminating gas, from the decomposed oil, and burns, when kindled, with a brisk flame. The oil at last acquires a treacle-like consistence and a brown colour. It is called *boiled oil*. *Printing ink* is made with boiled oil, lamp-black, and a little yellow soap. *Putty* is a mixture of linseed oil and whiting.

GROUP iii.—The Waxes and Spermaceti.

These substances differ from the glycerides inasmuch as their essential constituents are salts, not of the *triad* radical propenyl, but of some of the *monad* alcohol radicals homologous with methyl and ethyl (page 299). The acid radicals

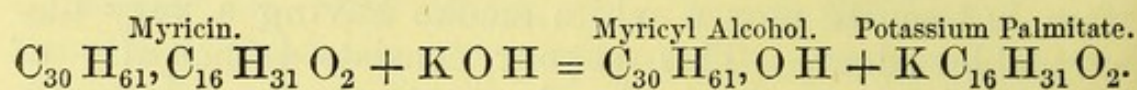
are those of the fatty acids, therefore they are homologues of the radicals of formic and acetic acids, so that the proximate constituents of these substances are true homologues of the formate or acetate of methyl or ethyl.

Spermaceti, found in cavities in the head of the sperm whale, consists chiefly of *cetin*; *Chinese wax*, obtained from the branches of certain trees in China, contains *cerin*, or *ceryl cerotate*; and *bees-wax* contains *cerin* and *myricin*, with two or more other constituents. Other vegetable waxes are known.

The following table will show the constitution of these proximate constituents, and their relationship to the simplest series of alcohols and fatty acids. Methyl formate, the lowest possible member of the series, is placed at the head for comparison :

		Methyl Formate.
	$C_2 H_4 O_2 =$	$CH_3, CHO_2.$
		Cetyl Palmitate.
Cetin,	$C_{30} H_{64} O_2 =$	$C_{16} H_{33}, C_{16} H_{31} O_2.$
		Ceryl Cerotate.
Chinese wax	$C_{54} H_{108} O_2 =$	$C_{27} H_{55}, C_{27} H_{53} O_2.$
		Myricyl Palmitate.
Myricin	$C_{46} H_{92} O_2 =$	$C_{30} H_{61}, C_{16} H_{31} O_2.$

These compounds are decomposed by caustic alkalies just as the glycerides are, the products being an alcohol, or hydrate of the basic radical, and a compound of the acid radical with potassium, or sodium :



This is, in truth, a process of saponification.

CHAPTER VIII.

VOLATILE OILS, RESINS, AND ALLIED BODIES.

THE compounds enumerated in this chapter are closely allied to one another. Many of them (the pure volatile oils, caoutchouc, &c.) are hydrocarbons; the rest are compounds of carbon, hydrogen, and oxygen. The constitution is not thoroughly understood, even of the best known members of the different groups. The phenomenon of *isomerism* (p. 295) is very common among them, many of the volatile oils having the same formula as that of oil of turpentine, $C_{10}H_{16}$. A few are derived from the animal kingdom (ambergris, castoreum, civet), but the great majority are the produce of plants. The following classification will be found useful.

GROUP I. *Volatile, or Essential Oils*.—Exist in several parts of many plants, either alone or mixed with resin, &c. They often separate by cold into two portions, one solid and camphor-like, which contains oxygen, and is called the *stearopten*; the other liquid, and free from oxygen, called the *elæopten*. They are commonly obtained by distilling the plant, or portion of the plant, with water. The oil comes over with the water and floats on its surface when it condenses. The liquid oils of lemon and orange peel, anise, cloves, lavender, thyme, and wormwood, are among those isomeric with oil of turpentine.

Whenever we perceive an odour in a plant, it is probable that a volatile oil is present, which gradually evaporates. But how incredibly diffused and diluted this must be in many plants may be inferred from the fact that scarcely a quarter of an ounce of volatile oil is contained in one hundred pounds of fresh roses, or orange-flowers. We most frequently find the volatile oils in the flowers and seeds, sometimes in the stalks and leaves, and

more rarely in the roots. The oils procured from the peel of certain fruits, as the lemon, citron, bergamot-pear, &c., are obtained by expression from the fresh rind.

An immense number of volatile oils are known. The following are among the most important.

a.) From the flower :

Oil of roses, a yellowish, thick fluid, with flakes resembling tallow floating in it.

Oil of orange-flowers (oil of neroli), colourless; becomes reddish on exposure to the light.

Oil of chamomile, a dark blue, thick liquid; becomes green, and finally brown, by age and light.

Oil of lavender, a yellowish, thin liquid.

Oil of cloves, yellowish, soon becomes brown; a somewhat thick fluid, heavier than water.

Oil of hops, greenish yellow.

b.) From seeds and fruits :

Oil of cumin, colourless; becomes yellowish, and finally brown, by age.

Oil of anise-seed, yellowish.

Oil of fennel, colourless or yellowish.

Oil of dill, yellow; becomes brown in the light.

Oil of nutmeg, a pale yellow, thin liquid; has the smell of nutmegs.

Oil of bitter almonds, yellow; heavier than water; consists chiefly of *benzoyl-hydride*, $C_7H_5O_2H$, but contains hydrocyanic acid, and consequently is very poisonous.

Oil of mustard, yellowish, of an extremely pungent smell, causing lachrymation; contains sulphur.

Oil of juniper, colourless.

Oil of savin, colourless, or yellowish; a thin fluid.

Oil of parsley, pale yellow; yields much stearopten.

Oil of lemons, from lemon-peel.

Oil of orange-peel, very like oil of lemons.

Oil of bergamot, from the rind of the bergamot orange, a pale yellow, very thin liquid.

c.) From the leaves and branches :

Oil of peppermint, colourless or yellowish, a very thin liquid.

Oil of balm, pale yellow, has an odour like that of lemons.

Oil of marjoram, yellowish or brownish.

Oil of thyme, when fresh, yellowish or greenish; when old, brownish-red.

Oil of sage, when fresh, yellowish or greenish; when old, brownish-red.

Oil of wormwood, dark green; soon becomes brown or yellow, and viscous in the light.

Oil of rosemary, colourless and very thin.

Cajeput oil, from the leaves of a tree growing in the Moluccas; the oil, when pure, is colourless; the crude oil is commonly green, and often contains camphor; it has a camphorated odour.

Oil of rue, pale yellow or greenish.

Oil of cinnamon, yellow; soon becomes brown in the air; heavier than water.

Oil of turpentine, the most common of the volatile oils, is contained in all our fir-trees, and exudes from them, mixed with resin, as turpentine. When purified, it is colourless and thin, and has an agreeable, penetrating odour. An ordinary sort, possessing a disagreeable, empyreumatic odour, obtained in the preparation of pitch from pine resin, is *crude oil of turpentine*.

d.) *From roots*:

Oil of acorns, yellow or brownish.

Oil of valerian, pale yellow or greenish; becomes rapidly brown and viscous on exposure to the air.

It is very remarkable, that we sometimes find several sorts of oil in one and the same plant. Thus, for example, we find in the orange-tree three different kinds of oil; one in the leaves, another in the blossom, and a third in the rind of the fruit.

GROUP II. *Resins*.—Solid bodies containing oxygen. They have faint acid properties. Common *colophony*, or *rosin*, which accompanies oil of turpentine, is a good example.

Other resins are *animé*, *copal*, *lac*, and *mastic*. *Amber* is a fossil resin.

GROUP III. *Balsams*.—Semi-solid masses. Some are mixtures of resins and volatile oils. Of these the most important are the different varieties of turpentine, including *Canada balsam*; *Burgundy pitch*; *balsam of Copaiba*, used in medicine, and *balsam of Mecca*, or *balm of Gilead*,

of which the finest kind is hardly known in this country ; others contain cinnamic acid (*balsam of Peru, storax, tolu*), or benzoic acid (*gum benzoin, dragon's blood*). A few curious and highly valuable animal substances, remarkable for their odour, may be included among the balsams ; namely, *ambergris*, found in the head of certain whales ; *castoreum*, obtained from the beaver ; and *civet*, from the civet cat.

GROUP IV. *Gum Resins* are mixtures containing resin gum and other ingredients. The most important are—

Asafoetida, from the root of a plant growing in Persia. Its disgusting odour is due to the presence of about three per cent. of a sulphuretted oil.

Gamboge, from the leaves and branches of a tree which grows in Cambogia and Ceylon.

Myrrh, from Arabia and Abyssinia. Imported in the form of reddish tears.

Olibanum, or incense. Burnt in religious ceremonies. Brought from Arabia.

Scammony, from Syria and Asia Minor. Aleppo scammony is the best. It has a strong odour and is used in medicine.

GROUP V. *Camphors*.—Solid volatile bodies of peculiar pungent odour. They contain carbon, hydrogen and oxygen. Common, or *laurel camphor*, $C_{10}H_{16}O$, obtained from *Laurus camphora*, a tree which grows in the East Indian islands, is the most important. *Borneo camphor*, $C_{10}H_{18}O$, is similar to common camphor. The stearoptens of some volatile oils are closely allied to the camphors.

GROUP VI. *Caoutchouc* and *Gutta percha*.—These substances, though so different in properties, are allied in composition to oil of turpentine. When pure they contain no oxygen, and are said to have the formula $C_{10}H_{16}$.

VOLATILE OILS.

Preparation of Volatile Oils.—*Experiment 1*.—Put one ounce of turpentine in a dish in a warm place, and when it has become liquid transfer it to a capacious flask ; pour upon it four ounces of water, and distil until about three-fourths of the water have passed over. Pour the residue, whilst still hot, into cold water, in which the non-volatile portion of the turpentine will soon congeal into a

solid mass (*resin*). A strongly smelling, colourless liquid—a volatile oil, commonly known under the name of *oil of turpentine*—floats on the surface of the water. Turpentine is the juice which exudes from pines and some other trees when the inner bark is cut through. We see from this experiment that it is a mixture of resin and volatile oil, which latter is carried over as vapour with the water and afterwards condensed.

If the distilled liquid is received in a pipette (p. 359), the lower end of which dips in water, the oil will collect in it and may easily be separated from the water.

Experiment 2.—Distil in the same manner half an ounce of cumin-seeds (which have been previously bruised in a mortar) in a retort containing four ounces of water, until two ounces of water have passed over. The drops floating upon the water consist of a volatile oil, *oil of cumin*; they have the smell and taste of the cumin-seeds, but in a stronger degree, while the residue remaining in the retort has scarcely the least smell or taste of them. All volatile oils possess a burning taste, and are somewhat rough to the touch; but the fat oils have a mild taste and an unctuous feel.

Experiment 3.—Pour a drop of some volatile oil upon a sheet of paper, and let it remain exposed to the air; the paper at first receives an apparent grease-spot, but this disappears after a time, because the oil gradually evaporates.

If the oiled paper is placed upon a warm stove, evaporation takes place much more rapidly. Aromatic oils are employed in this way for perfuming apartments. Usually a quantity of flowers, woods, and barks, finely cut up, are moistened with the oils, and scattered as a *fumigating powder* upon the stove.

Experiment 4.—Heat a quarter of an ounce of oil of turpentine in a vessel to boiling. A thermometer introduced into the liquid will indicate a temperature of about 328° F. (160° C.). Other oils boil with even more difficulty. The vapour may be inflamed by a taper, when it will burn with an intense sooty flame. It is easily extinguished by covering the vessel with a board; but *water must on no account be employed for extinguishing burning oils*. Then remove the oil from the fire. After it is cold,

mix it with some water, and again heat it; as long as any water is present the temperature of the fluid will not rise above 212° F. (100° C.). The ascending vapour is a mixture of the vapours of water and oil. The same thing occurs here as previously mentioned; the less volatile oil evaporates with the more easily volatile water. The oils remain unchanged at the boiling point of water, but at their own boiling point, between from 284° F. (140° C.) and 392° F. (200° C.), they become not unfrequently somewhat empyreumatic: this is the reason why water is always added in the preparation of oils, and also in the redistillation of them (*rectification*).

Experiment 5.—Inflame some drops of oil of turpentine put upon a shaving, and also a piece of camphor laid upon water; both bodies will ignite, and burn with a highly luminous and sooty flame. The volatile oils are far more easily combustible than the fat oils, which in order to burn with a flame must be heated to 662° F. (350° C.). We have consequently in oil of turpentine a convenient means for speedily lighting oil-lamps; it being merely necessary to smear the wick with a few drops of it.

Experiment 6.—Pour a mixture of half an ounce of absolute alcohol with half a drachm of oil of turpentine into a spirit-lamp; the mixture gives, when lighted, a strongly illuminating, but no longer a sooty flame, since all the carbon of the oil of turpentine is converted by the heat of the burning alcohol, rich in hydrogen, into illuminating gas, and then into carbonic anhydride and water. This mixture has sometimes been used in lamps constructed for the purpose, and which are so made that the liquid evaporates in them, and the vapour ignites as it issues from several small openings.

Volatile Oils and Water.—*Experiment 7.*—Drop some oil of cumin upon water; the oil floats on the surface without mixing with the water, for most of the volatile oils are *lighter* than water; but there are some, such as oil of cinnamon, oil of cloves, and oil of bitter almonds, which are heavier than water, and sink in it.

If the mixture is briskly shaken, the water becomes turbid, because the oil is thus divided into small, invisible globules, which are kept suspended in the water. The

water may be again clarified by filtration, but it retains the smell and taste of the oil, since a small quantity of it remains dissolved. Many such solutions are kept in the apothecaries' shops, under the name of *medicated* or *distilled waters*. It is well to keep them protected from the light, and in full vessels,—both light and air having a decomposing action on the volatile oils. They are commonly prepared by distilling with water the vegetable substance containing the oil, as thereby a more intimate combination of the water with the oil is effected than by merely shaking it up.

Volatile Oils and Alcohol.—*Experiment 8.*—Add a drop of oil of cumin to a little strong alcohol. It dissolves readily and entirely. All the volatile oils are soluble in alcohol, most of them even in alcohol of eighty per cent.; but the non-oxygenated oils, such as oil of turpentine, oil of lemons, &c., only in very strong alcohol. If an ounce of water, in which half an ounce of sugar has previously been dissolved, is added to the solution, we obtain cumin-cordial. In this manner, by the aid of various aromatic oils, the innumerable *cordials* or liqueurs occurring in commerce can be prepared. They were formerly manufactured from aromatic seeds, flowers, herbs, &c., by pouring brandy over them, the brandy being afterwards distilled or drawn off, whereby a spirituous solution of the volatile oils was likewise obtained.

Experiment 9.—If some drops of oil of bergamot, orange-flower, lavender, or rosemary, are dissolved in half an ounce of strong alcohol, we obtain a spirit of a very pleasant odour. In a similar way the innumerable kinds of perfumed waters are prepared, at the head of which stands the well-known *eau de Cologne*. The fumigating spirit also, which, instead of the fumigating powder, is often sprinkled on a warm stove, has a similar composition. Camphorated spirit, much used in medicine, is a solution of camphor in alcohol.

The volatile oils are not only dissolved by alcohol, but also by *ether* and concentrated *acetic acid*. A solution of oil of cloves, cinnamon, bergamot, and thyme, in acetic acid, is used as *aromatic vinegar*, on account of its refreshing odour.

The volatile oils may also be mixed with *fat oils*, and with some kinds of tallow and lard; hence by means of them an agreeable odour may be imparted to the latter, as, for instance, in hair oils, pomatum, &c., or grease-spots be dissolved and removed by them from various articles. Volatile oils mixed with alcohol yield, when shaken up with olive oil, a turbid, milky liquid, because the alcohol does not dissolve the olive oil; this behaviour may be taken advantage of for testing the purity of commercial oils.

Conversion of the Volatile Oils into Resins.—Experiment 10.—Let some oil of turpentine remain exposed to the air for some weeks, in a cup covered with paper, and afterwards put the cup in a warm place to evaporate the oil; it will not entirely volatilise, but will leave at first a viscous, and afterwards a vitreous residue. This residue is a kind of resin, obtained by the absorption of oxygen. Resins are often produced by the spontaneous oxidation of volatile oils. For this reason, *old* oil of turpentine is not suitable for removing grease spots, paint, &c., from clothing.

RESINS.

Preparation of the Resins.—Experiment 1.—Spread a little turpentine upon a board, and put the board for some time near a heated stove; the oil of turpentine evaporates, but the *resin* remains behind as an amorphous brittle mass. In some countries, incisions are made through the bark of fir-trees, and the turpentine which exudes is distilled with water, when the volatile oil passes over, the *resin* remains behind. Large quantities of such resin are now exported from the forests of America.

Resinous juices, which harden in the air, forming solid resin exude, either spontaneously or through incisions made for the purposes, not only from our fir-trees, but also from many other trees and shrubs, particularly those of hot climates. Almost all the resins occurring in commerce are procured in this manner.

Experiment 2.—Resin is deposited most abundantly in those parts of the trees where the branches join the trunk; wood impregnated with such resin is called resinous wood. If a piece of resinous wood is lighted at the upper end,

and held by a wire in an oblique position over a basin of water, one portion of the resin burns up with a sooty flame, while another part is melted by the heat, and runs down into the vessel beneath. *Resin is not soluble in water*; hence it hardens in the latter without mixing with it.

Fig. 107.



Experiment 3.—Pour strong alcohol upon some resinous wood, and let it remain for a day in a warm place; the resin is dissolved, and the woody fibre remains behind. The solution is poured into eight times its quantity of water, which is thereby rendered milky, because the resin is precipitated, but in such a state of fine division, that it floats about in the form of small globules. If this milky fluid is heated to the boiling point, the resinous particles soften and unite with each other in small lumps, which may be taken out and pressed together in larger masses. This is a third method of extracting resin from vegetable substances.

Resin and Water.—The resins as a general rule are insoluble in water, and therefore tasteless; but some of them in very small quantities may be dissolved, and these usually have a bitter taste. But many of the resins which occur in commerce contain some water in a state of minute division, and are thereby rendered dull and opaque; common pine-resin and boiled turpentine furnish examples of this.

Colophony or Rosin.—*Experiment 4.*—Heat a piece of solid turpentine, or else some pine-resin, in a spoon, till all the water is evaporated; the *anhydrous resin* will now appear perfectly transparent. In this state it is called *colophony*, or *rosin*, being *white* when it is moderately heated, but *brown* when the heat is so strong as to convert a part of the resin into *black pitch*. Colophony is so brittle, that it may easily be reduced to a powder. When the bow of a violin is rubbed with it, the rosin powder formed remains adherent to the hairs, and these then again adhere better to the strings of the violin. A similar effect is produced on the cords which sustain the weights in clocks when they are rubbed with rosin to prevent their

slipping. The resins, accordingly, exert an effect contrary to that of oil; by resin, a rough, uneven surface is produced; by oil, a smooth, slippery surface.

Action of Heat on Resins.—The experiment first performed reveals at the same time another property of resin; namely, its easy *fusibility*. Most of the resins require, in order to become fluid, a heat which is somewhat higher than that of boiling water. If melted rosin is poured upon a board, it spreads, and forms after hardening a solid, brilliant coating on the wood. The resins are hereby well adapted for protecting wood or metal from the penetration of air or water. For this reason, iron rails and iron ornaments are covered with a coating of pitch, to prevent them from being so quickly oxidised by the oxygen of the air; for the same reason, also, wine-casks and beer-barrels are smeared with pitch, that no air may penetrate into the casks, and that no beer may penetrate into the staves. The wood-work of ships, the hatches, &c., are covered with tar to keep out the sea-water and rain; and finally, also the solid and tenacious resin, *shell-lac*, which is formed by insects on the twigs of certain trees growing in the East Indies, is employed in the form of sealing-wax.

Sealing-wax.—*Experiment 5.*—Melt together in a small ladle one-fourth of an ounce of pale shell-lac, one drachm of turpentine, one drachm of cinnabar, and three-fourths of a drachm of prepared chalk; scrape out the mass while yet soft, and roll it out into sticks by the hands, moistened with water. The turpentine renders the sealing-wax more inflammable, and the cinnabar imparts to it the favourite red colour. Various other colours are given to it by chrome-yellow, ultramarine, Brunswick-green, lamp-black, bronze-powder, &c.

Rosin-gas.—*Experiment 6.*—When rosin is heated above its melting point, it *kindles* and *burns* with a luminous and sooty flame, leaving behind some charcoal. Therefore powdered rosin, when blown into a flame of the lamp, burns vividly.

Burnt Pitch.—If the rosin, after it has burnt for some time, is extinguished by putting a board over it, we shall have as a residuum a black, burnt resin, *ship-pitch*, and *cobbler's wax*, possessing great tenacity.

Lamp black.—*Experiment 7.*—If you hold a cone made of blotting-paper over burning pine-wood, it will soon become lined with soot. The well-known *lamp-black* is prepared on a large scale by a similar method. Resinous wood, or the resin itself, is burnt with an insufficient supply of air in a stove furnished with long flues, or with a chamber in which the smoke deposits its carbon on its passage through.

Experiment 8.—If some amber is scattered on glowing charcoal, a vapour having a pleasant balsamic odour is emitted from it as it smoulders away. Amber, frankincense, benzoin, and mastic are on this account frequently used for fumigating purposes.

Resin and Alcohol.—*Experiment 9.*—Wrap half an ounce of sandarach (a resin which comes from Barbary) in paper, and break it with a hammer into smaller pieces; then mix it with a drachm of sand, which has been previously freed from its finer particles by washing and afterwards thoroughly dried, and pour the mixture into a glass vessel, with two ounces of strong alcohol. Tie a piece of bladder over the vessel, or cork it, and let it remain for several days in a warm place, frequently stirring it round. The clear solution of resin thus obtained is called *lac-varnish*, because, when smeared over metal, wood, or paper, it leaves behind, after the alcohol has evaporated, a varnished, shining coat. If alcohol is poured upon the sandarach, unmixed with sand, resinous powder will cake together on the bottom of the vessel, forming a tenacious mass of resin, which dissolves much more slowly. To *lacquer* or *varnish*, then, is to coat the surface of anything with resin. By this coat of varnish, articles not only acquire a beautiful brilliancy, but are rendered at the same time impervious to air and water. When paper articles, as drawings, maps, &c., are to receive a coat of varnish, size or a solution of gum must previously be spread over them several times, as the solution of resin would otherwise penetrate into the fibres of the paper, and render it grey and transparent. This imbibition is usually prevented in wooden articles by smearing them with linseed oil before putting on the varnish. When the varnish is applied on places that are wet, white opaque spots are

formed, because the resin is separated by the water as a dull white powder.

Experiment 10.—Dissolve half an ounce of shell-lac in strong alcohol; a turbid liquid is obtained, as the shell-lac contains, besides the resin, small quantities of wax and mucilaginous substances, which float about undissolved in the solution of resin. This solution is also employed as a lac-varnish, but much more frequently as the so-called *French polish* of the cabinet-makers; that is, as a solution of resin, which they rub continuously upon the wood with a ball of linen, until the alcohol has evaporated. By this means a still smoother and finer polish is obtained than by merely applying the resinous solution with a brush, the marks of which frequently remain visible. The finer articles of furniture are usually French polished, the more common ones varnished.

Resins and Oils.—*Experiment 11.*—Mix half an ounce of dammara resin with some sand, and pour over the mixture two ounces of oil of turpentine; after a few days you will obtain an almost complete solution, for all the resins are soluble in the volatile oils. These solutions are also frequently employed as lac-varnishes; they dry, indeed, more slowly, but form a more tenacious coating, which is less liable to crack. The paler and finer varieties of varnish are principally prepared from amber, copal, dammara, shell-lac, sandarach, and mastic; the inferior and darker kinds, from amber-colophony, common colophony, turpentine, asphaltum, &c. A yellow colour is sometimes given to the pale varnishes by the addition of dragon's-blood, or gamboge.

The resins are likewise soluble in *fat oils*. Many of the ointments and plasters of the apothecaries consist of mixtures of fats and resins, and it is the latter which communicate to the former the property of adhering to the skin. Turpentine is sometimes employed for this purpose.

The use of resins in soap-making has already been noticed. They are also used by paper-makers, to prevent the porosity of the paper and the consequent running of the ink.

CAOUTCHOUC AND GUTTA-PERCHA.

Caoutchouc, or India-rubber.—When incisions are made in the bark of certain large trees, of which the *Siphonia* of South America and Java, and the *Ficus elastica* of Assam, are the most important, a milky juice exudes, which dries in the air to ordinary *caoutchouc*. It is often imported in the form of bottles, which are made by spreading the milky juice on clay moulds and drying it over a fire.

Experiment 1.—Caoutchouc at the ordinary temperature is hard and stiff, but it becomes soft when it is put into hot water or held near the fire. Cut from a piece of thin sheet caoutchouc, softened by heat, a square piece, apply it evenly round the ends of two glass tubes, and then clip off with a pair of scissors the ends of the strip in the direction marked out in the annexed figure: the fresh surfaces of the caoutchouc adhere firmly to each other (but still more closely when they are pressed together with the nail, yet without touching the freshly cut surfaces), and thus is formed a tube, which firmly tied at both ends, connects the two glass tubes air-tight to each other.

Fig. 108.

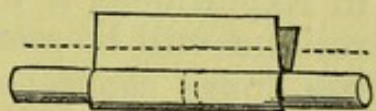


Fig. 109.



To the chemist, tubes made of caoutchouc are of the utmost value. They are sold of all sizes, and at a very moderate price. Caoutchouc corks also have been introduced into all laboratories. They are very useful, and may readily be bored with ordinary brass-tube cork-borers.

Caoutchouc is perfectly insoluble in water and alcohol, but pure ether, benzene (the lightest portion of coal-tar naphtha), and some volatile oils, including turpentine, dissolve it more or less perfectly. The best solvents are caoutchoucine (*vide infra*) and a mixture of carbon disulphide, with about 8 per cent. of absolute alcohol. When the solvent evaporates, the caoutchouc is left unaltered, and solutions of this kind are therefore largely used to render fabrics water-proof (the Mackintosh process). The gas bags used by chemists are made in this way.

Experiment 2.—Digest some india rubber in a stoppered bottle with one of the solvents named above. The solution so obtained may be used for waterproofing, and in the laboratory for a variety of useful purposes.

Caustic alkalies, cold or hot, have no action on caoutchouc. It is affected by strong nitric and sulphuric acids and by chlorine, but not by the majority of acids.

The *whole* of the caoutchouc does not dissolve in the above experiment, for caoutchouc appears to be a mixture of two substances, one of which is sparingly soluble.

Experiment 3.—Allow a little of the solution obtained in Experiment 2 to evaporate on a slip of glass. The residue will be seen to be *perfectly transparent and colourless* (pure caoutchouc), for the colour of ordinary india-rubber is due entirely to impurities.

Experiment 4.—Heat a little caoutchouc in an iron spoon over a lamp, very gently. It will melt to a viscid mass, which does not recover its elasticity on cooling.

Experiment 5.—*Marine Glue.*—Dissolve half-an-ounce of caoutchouc in two ounces of coal-tar with the assistance of heat, then add one ounce of shell-lac and heat the whole till it is homogeneous. A very powerful cement is obtained, which is readily melted by heat.

Experiment 6.—*Caoutchoucine.*—Distil a little caoutchouc very cautiously in a Florence flask with a large bent tube, keeping the receiver as cool as possible. A volatile oil of strong empyreumatic odour comes over, and can be rendered colourless by redistillation. This is called *caoutchoucine*. It is a mixture of two or more hydrocarbons, which are isomeric with one another and perhaps with caoutchouc itself. It is an excellent solvent for caoutchouc and resins.

Vulcanized Caoutchouc.—When caoutchouc is mixed with flowers of sulphur, or red antimonious sulphide, and the mixture kneaded in an iron vessel at a temperature of about 250° F. (121°C.), the mineral substance is taken up in a very curious manner. The caoutchouc acquires additional elasticity; it no longer hardens with cold, and may be heated to a much higher temperature than ordinary caoutchouc without becoming sticky. This is the so-called *vulcanized caoutchouc*, which is applied to so

many useful purposes. The kind made with antimony is called *red rubber*.

Ebonite.—When the mixture of caoutchouc and sulphur is heated for four hours to 302° F. (150° C.), a further change takes place, and the beautiful black substance called *ebonite* is obtained. At the temperature named it is soft, and can be moulded, but when cold it is as hard as horn, and can readily be cut and polished. Combs, beads, and many other articles are now made of it. It is of great value to the electrician, as it develops an amazing quantity of electricity when rubbed. It is insoluble, even in carbon disulphide.

Gutta-percha.—The juice of the *Isonandra gutta*, a tree which grows in Malacca and the islands of the Indian archipelago, is obtained like caoutchouc, which it resembles in many respects. It is imported in rough masses.

Gutta-percha dissolves easily, with the aid of heat, in benzene, oil of turpentine, carbon disulphide and chloroform. On allowing the solution to evaporate, the gutta remains as a brownish-red porous film.

Gutta differs in several important particulars from caoutchouc, with which however it appears to be identical in composition. At ordinary temperature, it is tough and flexible, but not elastic. At low temperature it is very hard and brittle, but at about 100° F. (38° C.) it becomes so soft that it can readily be welded or pressed into any desired form. It is a non-conductor of electricity, and is therefore largely used for coating telegraph wires.

When exposed to the air, gutta-percha undergoes slow oxidation and becomes brittle. A resin is formed. Vulcanized caoutchouc often undergoes a similar change.

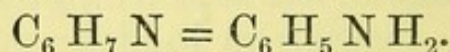
Gutta-percha is attacked by strong nitric and sulphuric acids. Hydrofluoric acid has no action on it, and it is therefore used for making bottles to contain that acid.

CHAPTER IX.

DERIVATIVES OF AMMONIA.—ALKALOIDS.

WE have already seen (p. 303) that a large number of compounds can be formed, which may be regarded as ammonia, in which the hydrogen is partly or entirely replaced by alcohol radicals (amines), or acid radicals (amides). The amines are true analogues of ammonia, and, like ammonia, combine with acids to form salts, which are called *compound ammonium salts*. The great majority of these artificial ammonia derivatives can only be prepared by difficult processes, and their description is therefore left to more advanced manuals of chemistry. The very important and interesting body called *aniline* will serve as a type of the amines.

ANILINE, OR PHENYLAMINE.

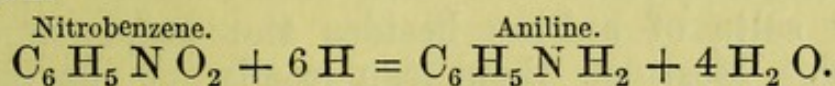


Experiment 1.—Take some powdered indigo; add to it about an equal quantity of very strong solution of potash, and distil in a retort as long as any brown oil comes over. Separate the oil from the ammoniacal water which accompanies it, and distil it again at as low a temperature as possible. The distillate is now colourless, and consists of aniline. This process is expensive, as it involves the loss of a good deal of indigo.

Experiment 2.—The hydro-carbon called *benzol*, or *benzene*, C_6H_6 , is now prepared in very large quantities from coal-tar. Take half an ounce of fuming nitric acid in a small beaker glass, or large test-tube, warm it slightly, and add, in very small quantities at a time, about an equal quantity of benzene. A curious change takes place, whereby the benzene is converted to *nitrobenzene*, $\text{C}_6\text{H}_5\text{NO}_2$

(p. 306). On adding a little water, the nitrobenzene falls to the bottom as a heavy yellow oil, insoluble in water, which possesses a smell closely resembling that of bitter almond oil. Wash the oil with water containing a little caustic soda and separate it by a pipette. Cover some granulated tin with strong hydrochloric acid and add the nitrobenzene a little at a time. Towards the end of the reaction the mixture should be heated, a new and peculiar smell is now observed. The liquid contains aniline combined with hydrochloric acid. Add excess of caustic soda which liberates aniline. Then shake the liquid with ether and decant the ethereal solution of aniline. If this is allowed to evaporate in a shallow basin, aniline is obtained.

The action consists in the replacement of oxygen by hydrogen.



Aniline is very little heavier than water (sp. gr. 1.03). It is sparingly soluble in water, readily in alcohol and ether. It boils at 363° F. (184° C.). Its aqueous solution is feebly alkaline. When exposed to the action of oxidizing agents, aniline yields a series of colours, which for brilliancy and variety are unrivalled. The manufacture of these so-called aniline, or *coal-tar* colours is now an important branch of industry. Their discovery is a remarkable example of the benefits which science confers on civilization and commerce.

Experiment 3.—Add a little solution of bleaching powder to a few drops of aniline on a white plate. A beautiful, though evanescent, violet colour will be produced. This is an excellent test for aniline.

Experiment 4.—When sulphuric acid is added to aniline, a crystalline mass of *sulphate of aniline*, $(\text{C}_6\text{H}_7\text{N})_2\text{H}_2\text{SO}_4$, is formed, which is very soluble in water, and may be purified by recrystallization.

Take a cold and somewhat dilute solution of this salt, and add to it a cold solution, also somewhat dilute, of potassium dichromate. Stir at intervals for ten or twelve hours. At the end of that time a black mass will have formed, which may be washed, first with water, and afterwards with

boiling benzene, which last dissolves out a tarry substance. The residue is the beautiful aniline purple, or *mauve*. It dissolves easily in alcohol, and the solution may be used for dyeing silk or woollen goods. Vegetable fabrics, such as cotton, do not hold the colour, but require the assistance of some *mordant*.

Experiment 5.—Add to a little aniline, in a test-tube, several grains of dry mercuric chloride, and boil the mixture for a few minutes. The liquid gradually becomes almost black; allow it to cool, and pour it into a large quantity of strong alcohol, which immediately acquires a magnificent crimson colour, and may be used for dyeing silk and wood. This is the colour generally known as *roseine*, or *magenta*. It is the hydrochlorate of a colourless base called *rosaniline*, $C_{20}H_{19}N_3.HCl$. The base may be thrown down by potash.

Many salts of aniline besides the sulphate can be prepared. Most of them can easily be purified by crystallization.

UREA.

Carbamide, or Carbonyl diamide, $COH_4N_2 = CO(NH_2)_2$.

This important body may be viewed as an amide, although it differs from most amides in that it combines with acids. Its constitution has already been explained (p. 305).

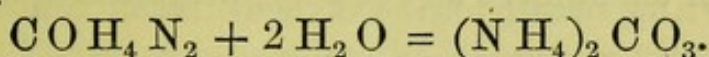
Experiment 1.—Evaporate a pint of fresh urine on a water bath* to about one-eighth of its bulk; allow it to cool. Reserve a small portion, and mix the rest with about an equal volume of a saturated solution of oxalic acid. A crystalline precipitate of *oxalate of urea*, $(COH_4N_2)_2H_2C_2O_4$, will form before long. When it has stood for a few hours, the liquid may be poured off, and the crystals pressed with blotting-paper, and dissolved in the smallest possible quantity of warm water. Chalk is now added in excess, when the oxalic acid is thrown down as the insoluble calcium oxalate (p. 356), and, together with the excess of chalk, may be removed by filtration. The clear liquid is

* A common saucepan without the cover makes a most excellent water-bath. It should be two-thirds filled with water, and the basin containing the substance to be heated rested on its mouth, the basin being of course somewhat wider than the saucepan.

now to be concentrated to a small bulk on the water-bath, when it will yield on cooling beautiful crystals of urea, slightly coloured, however, by the pigment of the urine.

Experiment 2.—The other portion of the concentrated urine may be mixed with strong nitric acid. Beautiful pearly crystals of *nitrate of urea*, $\text{COH}_4\text{N}_2\text{HNO}_3$, will separate.

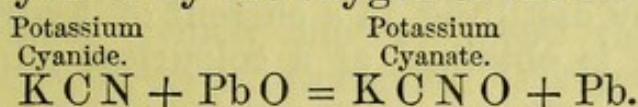
Experiment 3.—When urine putrefies, a kind of fermentation takes place, and the urea is converted to ammonium carbonate :



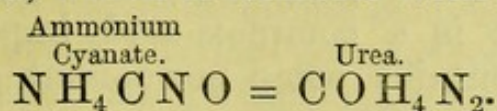
The smell of ammonia in putrid urine is well known.

Experiment 4.—*Synthesis of urea.*—Melt some potassium cyanide in a porcelain crucible, and add red lead, or litharge, in small portions at a time until it ceases to be reduced by the cyanide, and a little remains unchanged. Allow the crucible to cool, dissolve the mass in water, and add dry ammonium sulphate equal to twice the weight of the original cyanide. Evaporate to dryness on a water-bath, boil the dry mass with strong alcohol (methylated spirit), and filter. The filtered liquid will, on evaporation, yield colourless crystals of pure urea.

This process, discovered by Wohler, in 1828, is of great interest, because it afforded the first example of the synthesis of an organic compound (p. 102). Its details may be explained as follows:—The cyanide is first converted into cyanate by the oxygen of the lead-oxide :



When ammonium sulphate is added, ammonium cyanate and potassium sulphate are formed by double decomposition. But ammonium cyanate is instantly changed by heat into urea, which is isomeric with it :



NATURAL ALKALOIDS.

A large and important class of compounds containing nitrogen exists in different plants, to which the name of alkaloids, or natural bases, is applied. They are

commonly bitter in taste, and exceedingly active in properties, some of them being among the most virulent of known poisons. A few of them (nicotine, conine) are volatile liquids, containing only carbon, hydrogen and nitrogen, but the majority are crystalline solids, which contain oxygen as well. Many of them have an alkaline reaction, and all are capable of combining with acids and forming definite salts. There can be no doubt that the alkaloids belong to the class of ammonia derivatives, although few if any of them have as yet been prepared synthetically.

TOBACCO —NICOTINE.

Experiment.—Boil a quarter of a pound of common shag tobacco for a short time with a pint of water. Pour off the brown extract, and press the tobacco in calico. The liquid contains *nicotine* in combination with malic and citric acids. Evaporate it down on the water-bath until it has the consistence of oil; put it into a small stoppered bottle with about an equal quantity of strong potash solution and some ether or benzene, shaking the bottle at intervals for an hour. The potash liberates the nicotine from its combination, and the ether dissolves it. The ethereal solution may then be drawn off into a basin, and the ether allowed to vaporate, when nicotine in a somewhat impure state will remain. It may be purified by distillation, but as the boiling point is high (247° C. = 477° F.), a portion is destroyed by the heat. This may be avoided by distilling in a current of hydrogen, when the alkaloid is obtained as a colourless oil, of sp. gr. 1.011.

Nicotine has the formula $C_{10}H_{14}N_2$. It is a deadly poison, and possesses a terribly acrid and burning taste. It is strongly alkaline, and forms a definite series of salts. It is very soluble in water, alcohol, ether, and benzene.

Tobacco smoke is a complex and imperfectly known mixture. It contains some permanent gases, CO_2 , &c., and some empyreumatic oils, &c., in a state of vapour. The tobacco oil which collects in pipes, and which is often mistaken for nicotine, is an acid, tarry oil, analogous to that formed during the imperfect combustion of wood, &c. It contains some nicotine, but only a small quantity.

During the burning of tobacco, a slight deflagration is often noticed. This is due to nitre, which is always present in tobacco.

Conine is a volatile alkaloid, analogous in nicotine. It is contained in hemlock, and has the formula $C_8H_{17}N$.

PERUVIAN BARK—QUININE.

This celebrated bark, from which one of the most valuable of medicines is obtained, has sometimes been called Jesuits' bark, because its medicinal value was believed to have been first made known to Europeans by the Jesuits. It occurs on several species of the cinchona-tree, which flourishes in Peru, but is now cultivated with great success in India. It contains several alkaloids, of which quinine, $C_{20}H_{24}N_2O_2$, and cinchonine, $C_{20}H_{24}N_2O$, are the most important, in combination with a peculiar acid, called *quinic*, or *kinic* acid, $C_7H_{12}O_6$.

Experiment 1.—Into the percolating apparatus used in preparing tannin (p. 358), introduce some yellow cinchona bark (the yellow is the best kind) in fine powder, and pour on it some very dilute hydrochloric acid (1 of acid to 24 water). As the acid passes through, it combines with the alkaloids, and the liquid so obtained contains the hydrochlorates of quinine and cinchonine, together with a good deal of the colouring matter of the bark. The addition of acid may be repeated three or four times. To the acid solution, sodium carbonate is now added to alkaline reaction, which throws down the alkaloids, together with some colouring matter. After standing for some time, pour off the greater part of the liquid and evaporate the remainder, containing the alkaloids, to dryness on a water-bath. The dry mass may now be shaken up with ether in a small stoppered bottle. This takes up the quinine, which can be obtained as a resinous mass by evaporating the ether. After pouring off the ethereal solution of quinine, the residue may be digested with warm and strong alcohol. This takes up the cinchonine, and on allowing the alcohol to evaporate spontaneously, cinchonine separates in colourless anhydrous prisms. Neither alkaloid is quite pure when prepared by this process.

Experiment 2.—Both alkaloids dissolve very easily in dilute sulphuric and other acids. The substance generally sold as quinine is the neutral sulphate. Dilute sulphuric acid dissolves it easily. From all acid solutions the alkaloid may be precipitated by the addition of an alkali.

Experiment 3.—When quinine or any of its salts (acid solution of the sulphate answers well) is treated with chlorine water, and afterwards with solution of ammonia, a beautiful green colour is produced. This is the best test for quinine.

Experiment 4.—Commercial sulphate of quinine is sometimes adulterated with salicin (the bitter principle of willow bark). This may be detected by rubbing the dry salt with a drop of strong sulphuric acid. If salicin be present in any quantity, a bright red colour will be produced.

OPIUM—MORPHINE.

Opium is the dried juice of the unripe capsules of the white poppy, a plant which is extensively cultivated in the East. It contains many different alkaloids, together with a peculiar acid, *meconic acid*, $C_7H_4O_7$. The most important of the alkaloids is *morphine*, $C_{17}H_{19}NO_3$. The complete separation of the opium alkaloids is a matter of great difficulty, but the following experiment will illustrate the method.

Experiment 1.—Digest some opium, cut in thin slices, in lukewarm water for some time till a strong decoction is obtained; pour off the solution, and add just enough dilute ammonia to neutralize free acid. Evaporate on a water-bath to a small bulk, and then add excess of calcium chloride. *Calcium meconate* is precipitated while the alkaloids remain in solution as hydrochlorates. Filter and again concentrate, when more of the meconate will separate; after it has been filtered off, the hot solution will, on cooling, deposit crystals of the hydrochlorates of morphine and *codeine*. The first crystals may be dissolved in a little water and treated with ammonia, which throws down morphine. It is slightly coloured, but its solution may be purified by agitation with animal charcoal, which removes the colour.

Experiment 2.—From the calcium meconate formed in the above process, meconic acid may be prepared as follows: wash the calcium meconate and dissolve it in the smallest possible quantity of warm and very dilute hydrochloric acid. Add potash until a faint cloudiness appears, filter and concentrate to a small bulk. Somewhat coloured crystals of meconic acid will separate.

Experiment 3.—An aqueous solution of morphine or any of its salts gives, with neutral ferric chloride, a blue colour.

Experiment 4.—Put a minute particle of morphine, or one of its salts, on a white plate, and touch it with a drop of solution of iodic acid. Iodine will at once be liberated, and its presence is shown by its brown colour and the violet compound formed on the addition of starch paste. This is an excellent test for morphine.

Experiment 5.—Solution of meconic acid gives, with ferric salts, a *blood-red colour*. This test is so delicate that meconic acid can sometimes be detected where the quantity of opium present is too small to allow of the detection of morphine. The red colour is much like that given by iron salts with sulphocyanates, but the latter is instantly destroyed by solution of corrosive sublimate.

NUX VOMICA—STRYCHNINE.

The beans of *nux vomica* and *St. Ignatius*, and some other products of the order *strychnos*, contain two alkaloids, *strychnine* and *brucine*, both of which, especially the former, are deadly poisons. They cannot easily be prepared on a small scale.

Experiment.—Dissolve a minute crystal of strychnine on a white plate with a drop of strong sulphuric acid. Then rub with it an equally minute quantity of potassium dichromate. A beautiful blue colour, quickly fading to brick-red, is produced. Other oxidizing agents added to the sulphuric solution produce the same effect. This is an exceedingly delicate test for strychnine.

OTHER ALKALOIDS.

The following are some of the more important vegetable alkaloids, which have not been noticed in the preceding sections:

Aconitine, found in monkshood (*Aconitum napellus*). Said to be the most intensely poisonous of all the alkaloids.

Asparagine, in asparagus.

Atropine, in deadly nightshade (*Atropa belladonna*) and thorn-apple (*Datura stramonium*).

Berberine, in the barberry.

Caffeine, or *Theine*, in coffee and tea.

Colchicine, in colchicum.

Hyoscyamine, in henbane (*Hyoscyamus niger*).

Piperine, in pepper.

Ricinine, in castor seed (*Ricinus communis*).

Solanine, in the potatoe, &c.

Theobromine, in cacao beans.

Veratrine, in white hellebore (*Veratrum album*).

CHAPTER X.

COLOURING MATTERS.

OF the pigments used in painting, dyeing and colour-printing, a good many have already been described. Of *mineral* pigments the following are a few of the most important :

Reds.—Vermilion (HgS) ; red lead (Pb_3O_4) ; Venetian red (Fe_2O_3).

Yellows.—Chrome yellow (PbCrO_4) ; orpiment, or king's yellow (As_2S_3) ; yellow ochre (clay coloured by Fe_2O_3).

Greens.—Chrome green (Cr_2O_3) ; Schweinfurt, or imperial green (cupric aceto-arsenite).

Blues.—Ultramarine (a complex double silicate, containing a polysulphide of sodium) ; Prussian blue (ferric ferrocyanide) ; cobalt blue (a compound of cobalt oxide and alumina).

Among organic colours, some can only be prepared by artificial methods. To this class belong the splendid colours of almost every conceivable shade which are prepared from aniline and its congeners (p. 389). A beautiful crimson colour, called *murexide*, can be prepared from uric acid, which in its turn is obtained from urine or guano.

The term colouring matter is generally restricted to those colours which are obtained directly from animals and vegetables. To those we shall confine our attention here.

COLOURING MATTERS OF FLOWERS.

The beautiful and various tints of colour that we see in flowers are due to colouring matters which are very imperfectly understood. The reds and blues are generally soluble in water, the yellows only in alcohol. They are very evanescent, fading rapidly when exposed to light.

Experiment 1.—Digest some of the flowers of the blue iris, lobelia, or almost any blue flower, with a small quantity of warm alcohol. Pour off the alcohol and press the flowers in muslin. A blue solution will be obtained, which will yield a blue pigment when evaporated on the water-bath. Flowers of other colours will yield their colouring matters in the same way.

Experiment 2.—The colouring matter obtained in an impure state in the last experiment is called *cyanin*. Dissolve a little in water, and add to one portion of the solution a drop of acid. The blue colour changes to *red*. To another portion add a drop of soda. The blue changes to *green*. The red may be changed to green, and the green to red, by the addition of alkali or acid. Such a solution may therefore be used as a substitute for litmus. *Tincture of roses* is sometimes used for this purpose.

Experiment 3.—On adding a drop of solution of sulphurous acid to the aqueous solution the colour will first change to red, and then disappear. On standing, however, the acid will pass away and the colour return.

It is believed by some that the colour of most red flowers is due to cyanin reddened by an acid.

Experiment 4.—The red cabbage yields to water a colouring matter similar to, if not identical with, cyanin. The aqueous decoction yields a red with acids, a green with alkalies, and a purple with solution of alum; so that four different colours may be produced from the same solution.

Experiment 5.—Take the flowers of the orange-coloured *tropæolum majus*, or the brown calceolaria, and digest them with boiling water. A *purple* colouring matter will be dissolved out. Pour off the water, press the flowers, and digest them again with alcohol. A *yellow* colour will now be obtained and both may be separated from their solutions by evaporation. The flower is thus rendered white.

Some flowers therefore contain more than one colouring matter.

COLOUR OF LEAVES—CHLOROPHYLL.

Experiment 1.—Digest some fresh grass or other leaves with alcohol. A beautiful green solution is obtained,

which contains *chlorophyll*. In the solid state it is a dark green powder, which is insoluble in water, but soluble in alcohol and ether. Neither acids nor alkalies change its colour.

INDIGO.

A number of plants which grow in the East and West Indies contain an almost colourless substance called *indican*. When the leaves are macerated with water and allowed to ferment, the indican is decomposed and *indigo blue*, C_8H_5NO , is formed.

Indigo usually occurs in commerce in deep blue, friable cakes, which exhibit, when rubbed by the nail, a coppery colour and lustre. Its brilliant blue colouring matter is called *indigo-blue*; but besides this, the crude indigo contains other foreign substances, such as indigo-gluten, indigo-brown, indigo-red.

Indigo is quite insoluble in water, alcohol, ether, &c.; there is only one liquid known which can dissolve it, except in minute quantities, namely, concentrated sulphuric acid. The Nordhausen, or fuming acid, is the kind generally used, as six parts of it will dissolve as much as fifteen parts of the common acid. The indigo-blue chemically combines with the sulphuric acid, forming a blue compound soluble in water, which is called *sulphindigotic acid*. What we call tincture of indigo is principally a mixture of water, sulphindigotic acid, and free sulphuric acid.

The sulphindigotic acid combines with bases like a simple acid, forming salts. The best known of these salts is *sulphindigotate of potassium* (blue carmine), which is obtained as a deep blue precipitate when the sulphindigotic acid is neutralised by potassa. The blue carmine is indeed soluble in pure water, but not in water containing a salt in solution.

Indigo-white.—We can also, but in a very different way, render indigo soluble, by mixing it with bodies which, by taking oxygen from water, will impart hydrogen to it; for instance, with ferrous or stannous salts.

Experiment 1.—Triturate half a drachm of finely-powdered indigo with one and a half drachm of green

vitriol and two drachms of slaked lime; shake up the mixture in a four ounce bottle; then, having filled the bottle with water and closed it tightly, let it stand for several days; the indigo gradually loses its blue colour, and dissolves into a clean yellowish liquid. The body which effects the decolouration is the ferrous hydrate, which is separated by means of the lime from the green vitriol. The solution now contains the so-called *white*, or *reduced* indigo, C_8H_6NO . As soon as the clear liquid is exposed to the air, the extra atom of hydrogen is oxidised, and it again becomes blue. If you saturate a piece of blotting-paper or linen with the liquid, and then dry it in the air, it first becomes green, and then blue, and the blue colour formed adheres quite firmly, since it has not only settled *upon* but *in* the fibres of the paper. In dyeing establishments, such a solution of indigo is called the *cold vat*. A third method of rendering indigo soluble is by adding it, together with hot water, to a mixture of bran, woad, madder, &c., which (potassium and calcium carbonates being present) passes into fermentation. The fermentation is partly acid, and partly putrefactive; in both processes hydrogen is liberated and combines with the indigo. The colourless indigo dissolves in the alkaline liquid (*warm vat*).

By treating indigo with powerful reagents, such as nitric acid, chlorine, potash, &c., many interesting compounds can be obtained.

Experiment 2.—Heat a little indigo carefully in a watch-glass over which another watch-glass is inverted. A portion of the indigo will sublime in beautiful crystals.

Woad is a European plant, which likewise contains indican and yields indigo, but in far less quantities than the foreign indigo plants.

MADDER.

This important dye-stuff is the root of the *Rubia tinctorum*, a plant which is extensively cultivated in many parts of Europe.

In its natural state, the root appears to contain a glucoside, that is, a substance which is readily decomposed into glucose and something else. This com-

pound is called *rubian*. It is analogous to the indican of the indigo plant. It readily yields two colouring matters, *alizarin* and *purpurin*, the former of which is much the most valuable of the constituents of madder.

Experiment 1.—Digest some ground madder with warm water. A dark brown muddy liquid is obtained. This with dilute sulphuric acid yields a brown precipitate consisting of alizarin and purpurin with foreign bodies. The clear yellow liquid which remains contains nothing of importance. It may be removed by filtration, and the colouring matter washed with water. The precipitate is then to be boiled with concentrated solution of aluminium chloride (prepared by saturating strong hydrochloric acid with aluminium hydrate), which dissolves alizarin and purpurin but leaves a brown residue of impurities. On adding hydrochloric acid to the filtered liquid a red precipitate of the two colouring matters is obtained.

If it be desired to separate them, dissolve them, while moist, in dilute ammonia, and digest for some time with moist aluminium hydrate. They both combine with the metal, forming compounds called *lakes* (*vide infra*). The mixed lakes are then boiled with sodium carbonate, which dissolves out the *purpurin*. The residue is then boiled with hydrochloric acid, which leaves the yellowish-red *alizarin*.

When pure, alizarin and purpurin resemble one another closely. Both form yellowish-red crystals, which may to a great extent be sublimed unchanged. Both are sparingly soluble in water, more soluble in alcohol. Both are insoluble in acids, but easily soluble in alkalis, being thrown down unchanged when the alkali is neutralized. Both combine with the oxides of many metals (aluminium, iron, tin, lead &c.), forming insoluble compounds called lakes. It appears, however, that the colours produced by dyeing fabrics with madder are mainly compounds of alizarin.

Alizarin has the formula $C_{14}H_8O_4$. It is now prepared synthetically from anthracene (p. 410), and it appears probable that the artificial compound will supersede the root altogether. The formula for purpurin is $C_{14}H_8O_5$.

Madder lake.—A beautiful pigment bearing this name

is used by artists. The following is said to be the best receipt for it:

Experiment 2.—Enclose two ounces of the best ground madder in a strong calico bag capable of holding three times as much. Macerate the bag with one pint of cold soft water in a basin or mortar, pressing and twisting the bag as much as possible. Pour off the coloured liquid, and repeat the process until five pints of liquid are obtained. Nearly boil the liquid in an earthen vessel or large beaker, and pour it into another vessel containing 1 ounce of alum dissolved in 1 pint of boiling soft water. Stir and add $1\frac{1}{2}$ ounce of a saturated solution of potassium carbonate. When cool, pour off the liquid and add to the precipitated lake 1 quart of boiling soft water. When cold, pour off the washings and collect the lake, which should weigh half an ounce, on a filter. If required for water-colour painting it may be dried at a gentle heat and mixed with a little glycerine. Paints so prepared are called *moist colours*.

Dyeing with Madder.—Madder is a good example of what is sometimes called an *adjective*, as distinguished from a *substantive* colour. Substantive colours adhere to, or combine with the fibre of a fabric, and do not require any separate process to fix them. Adjective colours, on the contrary, have no power of fixing themselves to a fabric, but require the aid of some other substance, which, either by rendering them insoluble or by some other means, fixes them to the fabric. Such substances are called *mordants*. The following simple experiments illustrate their use in the case of madder:

Experiment 3.—*Madder red.*—Soak a piece of calico in the solution of impure aluminium acetate, which is obtained by mixing equal weights of alum and lead acetate (both in solution) and filtering from the insoluble lead sulphate. Let it hang in the air for a couple of days, and then soak it for two or three hours in a strong decoction of madder. It may then be removed and thoroughly washed, when it will be found to be dyed a fast and very permanent red.

Instead of soaking the calico in the mordant, the latter may be printed on in any required pattern (calico-printing).

The red dye will only appear on the parts where the mordant has been applied.

The action of the mordant here is obvious. The alizarin combines with the aluminium and forms the insoluble red madder lake. Aluminium acetate is called by dyers *red liquor*.

Experiment 4.—Various shades of violet and purple can be produced by using the mordant called *iron liquor* instead of red liquor. Iron liquor is prepared like red liquor, by mixing equal weights of green vitriol ($\text{FeSO}_4 \cdot 7\text{Aq}$) and lead acetate. It consists, of course, of ferrous acetate.

Experiment 5.—Mixtures of the above two mordants yield various shades of brown.

COCHINEAL.

A curious little insect, called the *cochineal insect* (*Coccus cacti*), is known, which lives exclusively on certain species of cactus. It was first discovered in Mexico in 1578, but is now cultivated in many tropical places. This insect contains about half its weight of a beautiful red colouring matter called *carmine*, or *carminic acid*. The female insects only are used. They are imported in the dry state.

Experiment 1.—Boil some powdered cochineal insects with soft water for a quarter of an hour. A splendid red solution will be obtained, for carminic acid is very soluble in water. It is equally soluble in alcohol, but almost insoluble in ether.

Experiment 2.—Add some moist aluminium hydrate to the above solution (cold). The alumina will combine with the colouring matter, and a beautiful crimson lake will be obtained.

Experiment 3.—To other portions of the solution add very small quantities of alum and stannic chloride respectively. The former will yield a crimson, the latter a fine scarlet colour. If these liquids are allowed to stand for a time, beautiful precipitates are deposited which are used by artists under the name of *carmine*. The precipitation appears to be due to the coagulation of the albuminoid matters of the cochineal.

Experiment 4.—The two liquids from which, in the last experiment, carmine was obtained may be cautiously

treated with sodium carbonate. The whole of the colour will now be thrown down as lakes, and, if excess of base has been avoided, they will possess brilliant colours.

Experiment 5.—Silk and wool may be dyed by immersing them in the solutions of Experiment 3. The colour adheres very strongly to the fabrics, but it slowly fades when exposed to light.

OTHER COLOURING MATTERS.

Brazil-wood, from the heart-wood of several trees growing in South America, imparts to different materials a beautiful but not very permanent red colour. It is employed also in the preparation of red ink, of drop-lake, &c. *Peach-wood*, *Nicaragua-wood*, *Sapan-wood*, &c., differ but slightly from *Brazil-wood*. (Colouring matter, *Brazilin*, crystallizes in orange-coloured needles, easily soluble in water.)

Safflower, the flowers of the dyer's saffron, is used for obtaining a brilliant rose-colour (for pink-saucers). (Colouring matter, *Carthamin*, soluble in water.)

The *alkanet-root* contains in its bark a resinous colouring matter, which is not soluble in water; silk is dyed violet with it, but alcohol, oils (as petroleum), and fats (as lip-salve), are coloured pink with it.

Sandal-wood (red sanders-wood), the rasped blood-red wood of a tree growing in the East Indies, contains likewise red, resinous colouring matter (*Santalin*).

The red dyes occurring in many *fruits*, as, for instance, cherries, raspberries, &c., are but slightly durable, and only used for colouring confectionery, cordials, &c.

Lac-dye is a red colouring matter very similar to cochineal, which is deposited with shell-lac (page 382) on the twigs of certain East Indian trees. The twigs with their coating are known as *stick-lac*.

Fustic is the rasped trunk-wood of a mulberry-tree which grows in the West Indies. It yields a yellow dye.

Quercitron, a nankeen-yellow powder, mixed with fibrous fragments, is obtained from the bark of the black oak, a tree of North America. (*Quercitron*, a yellow powder, soluble in water.)

Buckthorn, *Persian*, or *yellow berries*, the fruit of the

buckthorn and other species of *Rhamnus*, growing in warm countries. (*Rhamnin*, little known.)

Weld and *dyer's weed* are the names given to the *Reseda luteola*, dried after it has done blooming. *Luteolin*, crystallizes in yellow needles, soluble in water.)

The four last-mentioned colouring substances are principally used for dyeing silk, wool, cotton, and other materials, yellow.

Annotto occurs as a brownish-red paste, which is prepared from the pulp surrounding the seeds of the *Bixa Orellana*, and contains two colouring principles, a yellow and a red. The former is dissolved when the annotto is boiled with water, the latter on boiling it with a weak lye.

Turmeric, the root of a plant growing in the East Indies, is very rich in a resinous yellow dye, which is coloured brownish-red by alkalies. Paper stained with it may therefore be used like red litmus-paper for detecting alkalies. (*Curcumin*, an amorphous yellow mass.)

Saffron consists of the dried stigmas of the flowers of the *Crocus sativus*. Its application, in colouring articles of food and cordials yellow, is well known.

Logwood or *Campeachy-wood*, the reddish-brown interior wood of a tree of tropical America, is one of the most common colouring matters for dyeing blue, violet, and black. (*Hæmatoxylin*, in colourless crystals, which become speedily violet and blue in the air, owing to the ammonia always contained in the latter.)

Archil.—Several species of lichens, growing on the rocks in England and France, containing peculiar substances, which, although in themselves colourless, acquire a beautiful purple-red colour when they are acted upon by ammonia and air. It is common to putrefy the bruised lichen with urine, and then a red or violet-coloured paste is obtained (*cudbear*, *orchil*). By the addition of lime or potash, this red is changed into blue (*litmus*). We have examples of both these colouring matters in red and blue test-paper.

Experiments with the above Substances.

Experiment 1.—Take up some sandal-wood on the point of a knife and put it on a filter, and pour over it some alcohol; the alcohol which passes through has a red colour, and when poured upon a piece of wood, imparts to it an intense blood-red colour. Cabinet-makers frequently employ this solution for staining furniture. Alcohol acquires a pink colour when a small piece of *alkanet-root* is put into it. Water will not extract a red dye from either of these substances. Those colouring matters which are uncrystallizable, and soluble only in alcohol, are called *resinous*.

Experiment 2.—Boil for some time in a jar,—1st, either *fustic*, or *yellow berries*; 2nd, *Brazil-wood*; and 3rd, *log-wood*; each separately, with twelve times its amount of water; the decanted decoction of the first is yellow, of the second reddish-yellow, and of the third brownish-red; a sufficient proof that the colouring matters contained in these substances have been dissolved in the water. Dyers call these coloured decoctions *baths*.

Experiment 3.—Divide these colouring decoctions into two equal parts. Dissolve a quarter of an ounce of alum in one part of each, and then add to these a solution of potassium carbonate, as long as any precipitate subsides. Aluminium hydrate is precipitated; but with it the colouring matter, and hence the precipitates are coloured. These precipitates are called *lakes*. The lake obtained from the yellow berries occurs in commerce under the name of yellow lake, and that from Brazil-wood as drop-lake.

Experiment 4.—Prepare a solution of alum (*a*), another of stannous chloride (*b*), a third of ferrous sulphate (*c*), a fourth of potassium carbonate (*d*), a fifth of tartaric acid (*e*), and saturate a sheet of white blotting-paper with each solution. When dry, cut each sheet into three strips, smear one of the strips from each sheet, with the fustic, another of them with the Brazil-wood, and the third set with the logwood decoction, and again dry them. You will find that one and the same colouring matter produces a different colour, or shade of colour, upon each of the five

sheets. This colour will be very slight when the coloured decoctions are applied to mere blotting-paper (*f*). If you now immerse the coloured and dried slips in warm water, the colours will be for the most part dissolved from the three last tests (*d*, *e*, *f*), but not from the former (*a*, *b*, *c*), on which the metallic bases act as *mordants*.

These simple experiments give some idea of the interesting and complex arts of *dyeing* and *calico-printing*.

CHAPTER XI.

PRODUCTS OF DESTRUCTIVE DISTILLATION.

THE general nature of the interesting and important process of destructive distillation has already been explained (p. 188). We have seen that coal and many other non-volatile carbon compounds when heated in close vessels are decomposed; a portion of their carbon remains in the retort (coke, charcoal, bone-black, &c.), whilst a number of new and volatile bodies are formed which can be distilled over and collected. Of these, some at ordinary temperatures are *gases*, some *liquids*, and some *solids*. The manufacture of coal gas is a typical, and by far the most important example of destructive distillation. A reference to the table on page 189 will show how the products of the distillation of coal divide themselves into *gases* of various kinds which can be purified and collected, *ammoniacal water*, the chief source of the ammonium salts (p. 212), and *tar*, which collects below the water, and can easily be removed from it.

It is with the very complex and valuable liquid *tar*, and a class of substances allied to it, that we have to deal in this chapter.

COAL TAR.

A gallon of coal-tar may be procured for a mere trifle at any gas-works. It is black, viscid, like treacle, and has a very disagreeable smell. A portion of the ammoniacal liquor often floats on it and may be poured off and tested for ammonia, ammonium sulphide, carbonate, &c.

Distillation of Tar.—*Experiment 1.*—Introduce a pint of the tar into a retort made from a tin oil-can capable of holding a quart. The can must be fitted with a cork and bent tube similar to those shown on page 333. The tube should have an internal diameter of about half an inch,

and a length of about eighteen inches. Carry the end of the tube into a bottle to act as receiver, which at first must be kept very cold. Then apply a gentle heat to the retort. As soon as an ounce of liquid (naphtha) has distilled over, remove the receiver and replace it by another, which is no longer necessary to cool artificially. The heat must now be increased and the distillation continued until yellow scales begin to solidify in the long neck of the retort. A heavy oily liquid (dead oils) is obtained in this stage of the distillation. The receiver is now again changed and the temperature raised. The distillation may now be continued very carefully until rather more than one-fourth of the tar has been distilled off. In this last stage the distillate *solidifies* into a yellow, butter-like mass as it drops into the receiver. The neck of the retort may be warmed with a spirit-lamp from time to time to melt the solid that accumulates in it, and cause it to drop into the receiver. The black residue left in the retort must be poured out while still hot on to a stone or iron slab. When cold it appears as a hard, shining brittle mass; the *pitch*, or *asphalt*, now so largely used for coating woodwork and making pavements.

By this mode of conducting the distillation (fractional distillation) the tar is separated into four portions, namely, naphtha, or light oils, dead, or heavy oils, a solid portion, and pitch. Each of these is a very complex mixture. The constituents of these mixtures can only be separated on the large scale, and with great difficulty.

1. *Naphtha*.—*Coal-tar Naphtha*.—The portion which first comes over in the distillation of tar is very volatile, and is lighter than water. It consists chiefly of the hydrocarbon benzene, with its homologues. It contains also in minute quantity aniline and some other bases, and an *acid* portion, consisting of phenol or carbolic acid (p. 299), and its homologues. Coal-tar naphtha also contains a little naphthalene (p. 298).

The bases are first removed by agitating the naphtha with dilute sulphuric acid; the phenols are then removed by agitating with caustic soda, and finally the hydrocarbons, chiefly benzene and toluene, are removed by rectification nearly in the same manner as alcohol.

2. *Heavy oils.—Dead oils.*—The heavy oils of coal tar contain many hydrocarbons, some of which (notably naphthalene) are, when pure, solid at ordinary temperatures. But they are also valuable as containing the imperfectly acid bodies called *phenols* in considerable quantity. Ordinary phenol, or carbolic acid, is now prepared on a very large scale from this portion of coal-tar.

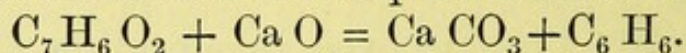
3. *Solid portion.*—This often amounts to a very large proportion of the whole products of distillation. The most important of its constituents are naphthalene and anthracene. From anthracene, *alizarin*, the chief colouring matter of madder, can be prepared.

4. *Pitch.—Gas-tar Asphalt.*—The nature of pitch is not yet well known. At a bright red heat, anthracene and other solid hydrocarbons can be distilled from it, and a coke remains which consists mainly of carbon.

PROXIMATE CONSTITUENTS OF COAL TAR.

Several of the constituents of coal tar are now important articles of commerce, and the student can therefore undertake a simple experimental study of them although he may not be able to prepare them from the tar.

Benzol or Benzene, C₆H₆.—Experiment 1.—Distil a mixture of one part benzoic acid and three parts slaked lime. Benzene and water pass over. The former has a specific gravity of only 0.88, and therefore floats on the water. Add some pieces of caustic potash to the mixture to absorb traces of acid. After standing some time, decant the benzene into a small flask containing solid calcium chloride. Cork and allow the mixture to stand twenty-four hours; then distil off the perfectly pure benzene by means of a water bath and bent tube, taking care to keep the receiver very cold and to prevent the vapour from coming near a light. The formula which describes the formation of benzene in this experiment is as follows:



Commercial benzene is obtained as above described from gas tar. It is never pure, but contains higher hydrocarbons, of which the most important is *toluene*, C₇H₈, the homologue next above benzene, and also

a curious compound called *thiophene* C_4H_4S . Nevertheless, when carefully rectified from a water bath, the less volatile portions being rejected, it will do very well for the following experiments:

Pure benzene boils at $80^\circ C.$ ($176^\circ F.$). It has a peculiar but not unpleasant odour.

Experiment 2.—Benzene ignites very easily when a light is applied, and burns with a bright smoky flame.

Experiment 3.—Benzene evaporates very easily at ordinary temperatures, and the vapour is very inflammable. Shake a few drops of the liquid in a small bottle of oxygen and apply a light. An explosion will be produced.

Experiment 4.—Put some cotton wool into a tube drawn to a jet at one end. Pour a few drops of benzene on the cotton, blow gently through the tube, and apply a light to the jet. The mixture of air and vapour will burn with a luminous flame.

Experiment 5.—Benzene is a powerful solvent of oils and fats. Shake a known weight of oil with benzene in a test tube. It will dissolve in considerable quantity. On allowing the solvent to evaporate on a basin the oil will be recovered unchanged. If the basin is heated on a water bath for some time and again weighed, the increase of weight will correspond with that of the oil employed. This experiment shows how benzene may be employed in estimating the quantity of oil or fat present in any substance. It also illustrates the use of benzene, in removing grease-stains from articles of clothing.

Experiment 6.—Immerse a beaker or tumbler containing benzene in a mixture of snow or crushed ice and salt. The benzene crystallizes and may be separated from other substances by pressing it rapidly in a cloth. Toluene does not solidify at $-20^\circ C.$ ($-4^\circ F.$). At $5^\circ C.$ ($41^\circ F.$), the benzene, now very pure, once more becomes liquid.

The conversions of benzene into *nitrobenzene*, and of nitrobenzene into *aniline*, have already been explained (p. 389). Pure aniline does not yield the beautiful aniline colours, but only a mixture of aniline and toluidine. *Toluidine*, $C_7H_7NH_2$, is a homologue of aniline, prepared from toluene as aniline is from benzene.

Phenol, or Carbolic Acid, $C_6H_6O = C_6H_5OH$.—This valuable and interesting compound is now prepared on a large scale from coal tar, and may be purchased in the pure state at a cheap rate. When pure it crystallizes in colourless deliquescent needles, which melt at $41^\circ C.$ ($106^\circ F.$) and boil at $180^\circ C.$ ($356^\circ F.$). It has a peculiar tar-like odour, has a specific gravity of 1.056 and is very poisonous.

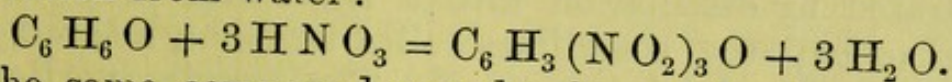
Experiment 1.—If the heavy oil of coal tar is distilled by itself, and the portion which comes over between 150° and $200^\circ C.$ (302° and $392^\circ F.$) is collected by itself, this portion will contain the great bulk of the phenol of the tar. When a concentrated solution of potash and a little solid potash in powder is added to it, the liquid becomes semi-solid from the formation of the white crystalline *potassium carbolate*. This is dissolved in hot water, the neutral oil removed from the surface, and the salt neutralised with hydrochloric acid. Impure phenol rises to the surface as an oil, and may be washed, dried by agitation with calcium chloride, and purified by repeated distillations and crystallizations. The purification is, however, a difficult process, and for many purposes is not necessary.

Experiment 2.—Dissolve some pure phenol in water, and observe that it is sparingly soluble in cold, but somewhat more in hot water. It dissolves easily in alcohol and ether. The solutions, when pure, have a neutral reaction, although phenol combines with bases, and so far has an acid character. The salts are, however, somewhat indefinite. With bromine water, even very weak solutions of phenol give a white precipitate of *tribromophenol* $C_6H_3Br_3O$.

Experiment 3.—Phenol and its homologues are among the most powerful of *antiseptics*. Suspend a piece of fresh meat above a little phenol in a bottle. The vapour of the phenol will entirely prevent the putrefaction of the meat, which will in time dry up into a horny mass. For this reason phenol is often used in preserving anatomical specimens. Its action in *disinfecting*, which is its most important use, is probably due to its power of destroying the organized germs which appear to be the active agents

of contagion in some zymotic diseases. It has, however, very little power as a deodorizer.

Trinitrophenol; Picric, or Carbazotic Acid, $C_6H_3(NO_2)_3O$.—*Experiment 1.*—Add fuming nitric acid, a drop at a time, to a small quantity of phenol in a large test tube. Violent action takes place. When this begins to diminish, the mixture must be warmed and ultimately boiled, for otherwise mono- and dinitrophenic acid will be formed. On concentrating the acid liquid, beautiful yellow crystals of trinitrophenic, or, as it is more often called, picric acid are obtained, and may be purified by recrystallization from water:



The same compound may be formed by the action of nitric acid on indigo, silk, and other substances. The dyeing of animal substances by nitric acid (p. 162) is probably due to the formation of this compound.

Experiment 2.—Picric acid is sparingly soluble in cold water, but the solution, even when it contains only 1 part acid in 10,000 of water, has an intense yellow colour. It is very soluble in alcohol. The solutions are used for imparting a beautiful and permanent colour to silk.

Experiment 3.—Neutralize a warm and moderately concentrated solution of sodium carbonate with picric acid. *Sodium picrate* separates in yellow crystals as the solution cools. Other picrates may be formed by double decomposition with the sodium salt. *Potassium picrate* requires 260 parts of cold water ($15^\circ C. = 59^\circ F.$) for solution, and for this reason sodium picrate, which is very soluble in water, is sometimes used as a test for potassium.

Experiment 4.—Heat a little potassium picrate on a knife blade. It burns with explosion. In a close vessel the explosion is violent, and the sodium salt has even been proposed as a substitute for gunpowder.

PARAFFIN—PARAFFIN OIL—PETROLEUM.

When coal, peat, lignite, bituminous shales, &c., are exposed to destructive distillation at a temperature which is not allowed to rise above low redness, very little gas is obtained, but instead, a considerable quantity of tar or oil, which differs in composition from the tar obtained in

manufacture of gas. This tar is called *crude paraffin oil*, because it contains in solution the solid hydrocarbons which are grouped together under the name of *paraffin* (p. 298). A kind of coal, called *boghead coal*, found at Bathgate near Edinburgh, is a good material for the preparation of paraffin oil. It is said to yield as much as 100 gallons of crude oil per ton.

The crude oil is purified in much the same manner as gas tar. It is freed from basic and acid constituents by treatment with sulphuric acid and caustic soda. The remaining oil consists entirely of hydrocarbons, and it is separated by distillation into a light and very volatile *spirit*, or naphtha, sometimes called *benzoline*, a *burning oil*, and a heavy, or *lubricating oil*. The heavier portions of the oil deposit solid *paraffin* on cooling, and this paraffin is purified and applied to the manufacture of candles and other useful purposes. After the paraffin has been separated, the oil is again distilled with other portions of the oil, and a large proportion of oil, suitable for burning in lamps, is thereby obtained. The lighter portions of the oil cannot be used with safety in ordinary lamps, as the vapour, mixed with air, is very explosive. Terrible accidents often occur from the use in lamps of oil which contains too large a proportion of the lighter constituents. The safety of paraffin or petroleum oil may be tested in the following manner:

Experiment.—Fill a beaker two-thirds full of the oil and cover it with a card, and immerse it in boiling water till the temperature of the oil rises to 100° F. (38° C.), taking care to avoid draughts of air, which might blow away the vapour. Then bring a taper very near to the surface of the oil. If a flash of flame appears the oil is unsafe, and ought on no account to be used. In short, the *flash point* of the oil ought to be above 100° F.

Pure paraffin oil is a very complex mixture of hydrocarbons, the greater number of which belong to the saturated, or marsh gas series, C_nH_{2n+2} .

Petroleum.—In many parts of the world, as, for instance in North America, Burmah, Persia, and the West Indies, hydrocarbon oils, known variously as *petroleum*, *rock oil*, *mineral tar*, or *naphtha*, &c., are found in great abundance

below the soil, and can be obtained by sinking wells. These oils are very similar to paraffin oil, and some of them yield a great deal of solid paraffin. Rangoon tar yields as much as eleven per cent. of this substance. The petroleum of Pennsylvania has been found to yield at least twelve homologous liquid hydrocarbons, besides gases and solid paraffins. They all appear to belong to the C_nH_{2n+2} series.

Asphalt.—Compact Bitumen.—A solid black substance, between coal and petroleum in composition and properties. It melts at about $100^{\circ}C.$, and burns with a smoky flame. It is found in considerable quantities in some parts of the world, notably on the shores of the Dead Sea and in the celebrated bitumen lake—one mile and a half in circumference—of Trinidad. There are also small deposits of asphalt in Cornwall, Derbyshire, and other parts of England.

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