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MANUAL OF CHEMISTRY FOR DISPENSERS

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By Command of the Army Council,

H. J. Creedy

THE WAR OFFICE,
30th September, 1931.

INTRODUCTION

The contents of this book comprise subjects C and D of Part A, of the course of instruction for the qualification of dispenser in the Royal Army Medical Corps.

The compilers have endeavoured to include such general practical knowledge of elementary chemistry and physics as will enable a student dispenser to understand the main principles of pharmacy and of the first-aid treatment of poisons.

This information has of necessity been condensed in such a form that the whole syllabus can be taught in the training period prescribed in Standing Orders for the Royal Army Medical Corps.

Each chapter or lesson concludes with a series of simple experiments, described and illustrated; a few questions relative to the contents of the chapter are also appended.

The book consists of the following :—

A detailed list of contents.
Part I.—Inorganic Chemistry.—Chapters 1 to 24.
Part II.—Organic Chemistry.—Chapters 1 to 8.
A list of definitions.
Appendix 1. A list of apparatus and chemicals required for experiments, and the sources from which these may be obtained.
Appendix 2. Table of Atomic Weights.
Appendix 3. Table of Solubilities.
Appendix 4. Normal and Standard Solutions.
Appendix 5. Table of Incompatibilities.

The appendices do not form part of the Syllabus for Part " A " of the Dispensers' Examination, and are included for reference when required.

No index has been compiled, as it is considered that the headings of the chapters give all the information necessary for ready reference.

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PART I

INORGANIC CHEMISTRY

CHAPTER 1

MATTER AND ITS COMPOSITION

Chemistry is the study of the composition and the properties of *matter*, and is divided into two main sections, namely, *organic chemistry*, which deals with animal and vegetable matter, and *inorganic chemistry*, which treats of mineral matter.

Matter and Molecules

Matter is the name given to all substances existing in the world, and is found in three distinct forms—as a solid, a liquid, or a gas. Whatever form it assumes, matter is composed of a large number of minute particles, and to these particles the name *molecules*, or "little masses," has been given. All molecules of one substance are alike—for example, all the molecules of a lump of pure gold are composed of gold and all the molecules of a drop of water consist of water. Molecules are extremely small, invisible to the naked eye, and are in a state of perpetual motion. This motion or movement of the molecules determines the form the substance takes.

Cohesion

If the movement is very slight and the molecules are packed closely together, the substance will be a solid ; if, however, the movement is considerable and the molecules are widely separated, then the substance will be a gas. The power that keeps the molecules together is *cohesion*. The closer they stick together the harder or more solid the substance is. The molecules of gas have no cohesion, and therefore they become widely dispersed when allowed to escape into space. Liquids hold a position mid-way between solids and gases. In them the cohesion is sufficient to keep the molecules in contact, yet allowing them to slip or glide over one another. To illustrate the difference, suppose the molecules of a solid are represented by bricks cemented together to form a wall, the bricks cannot be separated unless considerable force is used, the cohesion of the bricks is so great that the wall appears as a solid mass. Now let the molecules of water be represented by the grains of sand in a cart. If the cart is tipped up the sand will slide out, one grain sliding over the surface of another, sufficient cohesion being present to keep the particles in contact.

The molecules of gas could be represented by numerous small balloons filled with hydrogen. To keep them together one would require to enclose them in a box, otherwise, if allowed freedom, they would float away from one another and disperse in the air.

The definition of a molecule is the smallest particle of matter that can exist by itself.

Atoms and Chemical Change

Molecules possess a structure, for they themselves are built up of smaller particles known as *atoms*, and the force which holds these atoms together is called *chemical affinity*. A molecule of water is composed of two atoms of hydrogen and one atom of oxygen, both of which substances are gases; the force which prevents these gases from separating is chemical affinity. When, for some reason, separation takes place, for example by passing an electric current through the water, then a *chemical change* or *action* occurs. Heat is generally produced during a chemical action.

An atom is defined as the smallest particle of matter which can take part in a chemical change and which cannot be further divided.

Physical Change

There is another kind of change that the molecule may undergo. For example, water when boiled is converted into steam, that is, the water has altered its state from a liquid to a gas. The molecules which comprise steam are still composed of a union of hydrogen and oxygen, therefore no chemical change can have taken place. The alteration is purely one of form and is spoken of as a *physical change*.

Elements

When the molecules consist of atoms identical to one another in every respect, then the molecules are said to be *elementary molecules*, and the substances composed of these molecules are called *elements*. Taking gold again as an example, all the atoms of the gold molecule are the same, and therefore gold is classified as one of the elements. There are over eighty of these elements known to exist. This is a very small number compared with the millions of other substances which go to form the land, the sea, and the air, and all the vegetable and animal life that exists in the world.

Compounds

What are these other substances? They are called *compounds*, that is, substances the molecules of which are composed of two or more dissimilar atoms. Take for example salt and sugar. They are both compounds. The molecule of salt is formed by the union of an atom of sodium (Na) with an atom of chlorine (Cl); the molecule of sugar, on the other hand, is built up of three different types of atoms—carbon (C), hydrogen (H), and oxygen (O).

States or Forms of Matter

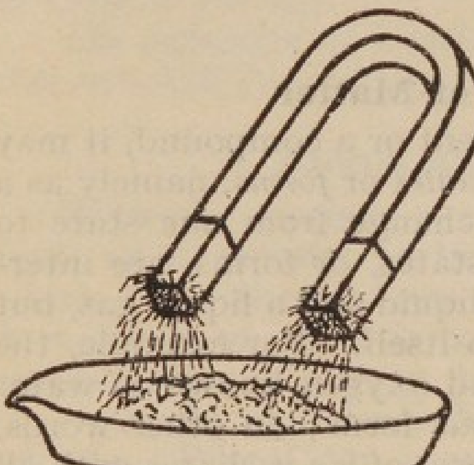
Whether a substance be an element or a compound, it may show itself in one or other of three *states* or *forms*, namely as a *solid*, a *liquid*, or as a *gas*. The change from one state to another is a physical one. These states, or forms, are interchangeable. A solid can become a liquid and a liquid gas, but each possesses a quality peculiar to itself. For example, the chemical compound of hydrogen and oxygen known as water may exist in solid, liquid, or gaseous form; in other words, as ice, water, or steam. If in the form of ice it shares with all other solids the quality of retaining its shape when placed in a container, whereas if in the liquid state it conforms to the character of all liquids, namely, that when poured into a container it will take the shape of that part it fills, and finally, if in the form of steam, it will not only fill the container completely but will, unless prevented, escape and diffuse into the air. Taking a concrete example, a piece of ice dropped into a bottle retains its shape so long as it keeps frozen; if, however, the ice melts, the water formed will assume the shape of that part of the bottle it fills. Now if sufficient heat is applied to boil the water, then the whole bottle will be filled with steam, which will force its way out through the mouth and disperse.

Mixture

There is still another term that requires explaining, namely, the term *mixture*. *This is the name given to substances which are formed by the blending together or mixing of two or more elements or compounds, so that each element or compound retains its distinctive character and properties.* No chemical change takes place to alter the ingredients, which can often be separated again by mechanical means. For example, iron filings, sand, and sugar when added to one another form a mixture. The iron filings and sugar can still be distinguished from the grains of sand in the mixture. By means of a magnet all the filings can be removed, and the sugar being soluble can be separated from the sand by adding water, shaking thoroughly, and straining through cloth. The sand can be removed from the cloth after drying. The water containing the sugar in solution is then poured off into a separate dish and allowed to evaporate, when the sugar will be recovered. The iron filings, the sugar, and the sand have suffered no alteration either in their character or properties by their contact with one another in the mixture.

PRACTICAL CHEMISTRY—CHAPTER 1

A Mixture



Of Solids

Expt. 1

Mix in a tin or crucible a teaspoonful of iron filings and flowers of sulphur.

By means of a magnet show that the iron filings can be removed from the sulphur.

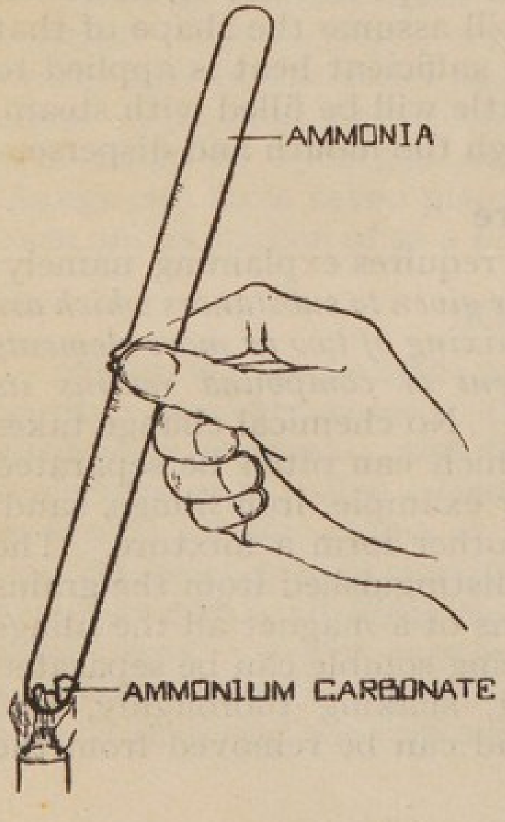
Of Liquids

Expt. 2

Add some ether to a little olive oil in a test tube and mix thoroughly.

Evaporate the ether by applying gentle heat, and the olive oil will be left unchanged.

Of Gases



Expt. 3

Heat a pinch of ammonium carbonate in a test tube and allow the ammonia gas evolved to mix with the air in another test tube held upside down over the mouth of the first test tube.

Show the presence of ammonia in the empty test tube by introducing a moistened piece of red litmus paper, which turns blue.

Add a little water and shake thoroughly and again test for ammonia, taking care not to touch the side of the tube with the litmus paper.

The ammonia dissolves from its mixture with air, which is left unaltered.

A Compound

Of Solids

Expt. 4

Mix some sulphur and very fine iron filings and heat in an old spoon over a flame.

The mixture changes colour. A chemical action takes place resulting in the formation of iron sulphide.

The iron cannot now be separated from its combination with sulphur by means of a magnet.

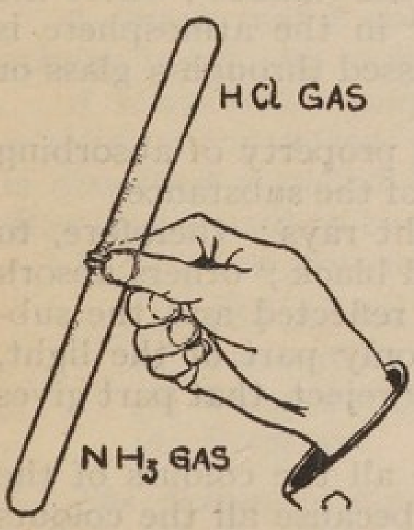
Of Liquids

Expt. 5

Add a solution of magnesium sulphate to a solution of sodium carbonate.

A white precipitate of an entirely new substance results from the chemical action which has taken place.

Of Gases



Expt. 6

Wet the inside of two test tubes, the first by adding one-quarter of an inch of pure hydrochloric acid ; and the second by adding the same amount of liquor ammonia fortis and then emptying the tubes.

Place the two test tubes mouth to mouth as in the diagram and note the immediate formation of a solid substance, ammonium chloride, where the two gases come in contact.

QUESTIONS

Question 1.—State the difference between a mixture and a compound.

Give examples of liquid, gaseous and solid mixtures, and compounds.

Question 2.—What is meant by the terms Physical Change and Chemical Change ? Give an example of each.

CHAPTER 2

GENERAL PROPERTIES OF MATTER

A SUBSTANCE possesses certain properties by which it can be identified. The most obvious of these are detected by man's special senses. *Colour* is noted by the eye, *taste* by the tongue, and *smell* by the nose.

Colour

Some substances are coloured, some are not. To explain the reason why they differ requires some knowledge of what *light* is. Light is composed of the three *primary colours*, Blue, Yellow, and Red, and where they merge into one another the secondary colours appear. Where blue changes to yellow shades of green are visible, where yellow changes to red there is orange, and where red meets blue, violet and indigo are seen.

The complete display of colours, violet, indigo, blue, green, yellow, orange, and red, is known as the *spectrum*. These colours or spectrum become evident in the rainbow, where the sunlight striking small drops of water in the atmosphere is broken up, or when rays of light are passed through a glass or crystal prism.

Matter may or may not possess the property of absorbing light, it entirely depends on the nature of the substance.

Some substances absorb all the light rays ; therefore, to the eye they appear dark and are called black ; others absorb no light at all, the light so to speak is reflected and the substance appears white ; others absorb only part of the light, and whatever part of the spectrum they reject, that part gives colour to the substance.

For example, coal is black because all the colours of the spectrum are absorbed, a lily is white because all the colours are reflected, a rose is red because all the colours except red are absorbed, and its leaf green because the blue and yellow rays are returned and combine to form green.

Taste

Taste may be closely allied to smell, for many substances on entering the mouth not only excite the taste buds on the tongue but also the olfactory nerve endings in the nose.

Tastes may be classified as *bitter*, *sweet*, *acid*, and *saline*. Examples of bitter substances are hops and quinine ; *sweet*, sugar and glycerine; *acid*, vinegar and lemon juice; and *saline*, salt and sea water.

There are many substances without any taste, and they are invariably insoluble—for example, glass and china.

Smell

This property possessed by certain substances may be *aromatic* or *etheral*—for example, the scent of flowers and perfumes; *pungent* or *sharp*, as in the case of tobacco smoke and burning sulphur; *nauseous* or *fœtid*, as, for example, decomposing animal or vegetable matter.

Many substances have no smell at all, such as air, water, and salt; whilst others are so distinctive in their smell that they can be identified by this property alone. For example, no one could mistake the smell of an orange or of coal gas.

Solubility

There are other properties possessed by matter in addition to those of colour, taste, and smell, one of the most important of which is that of *solubility*.

A great many substances possess the property of dissolving in liquids, to form solutions.

A solution is a uniform mixture of either a gas, a liquid, or a solid with a liquid. The liquid in which any one of these is dissolved is called the *solvent* and the substance which is dissolved is said to be *soluble* in that liquid.

The factors governing solubility are as follows:

(1) The nature of the substance.
(2) The nature of the solvent.
(3) Temperature.
(4) Pressure.

The Nature of the Substance.—Some substances are more soluble than others. In a small amount of water a large quantity of sugar can be dissolved, whereas the minutest particle of sand dropped into the ocean would remain undissolved.

This also applies to liquids. Methylated spirits and water are soluble in one another, but paraffin will not mix with water, and forms a film on the surface; it is therefore insoluble in water.

The Nature of the Solvent.—Substances may be soluble in one liquid but insoluble in another. For example, sugar will dissolve in water but will not do so in paraffin oil; butter, on the other hand, will dissolve in paraffin oil but remains insoluble in water.

Temperature.—The higher the temperature of the liquid, the greater will be the amount of the substance dissolved. Boiling water dissolves twelve times more saltpetre than cold water. There is one exception to this; gases are all expelled from solution on boiling. As water approaches boiling point, bubbles of gas will be seen to rise to the top, and before boiling point is reached no gases will be left in solution. On the other hand, water near freezing point will dissolve twice as much chlorine gas as it does at body temperature.

Pressure.—Pressure does not affect the solution of solids or liquids, but it has a marked effect on the solution of gases. Soda water is a good example of this. Carbonic acid gas is forced into the water under pressure and remains in solution so long as the bottle is closed; when, however, it is opened and the pressure relieved, then bubbles of the gas are given off. A gas can be completely removed from solution by placing the liquid containing it in a vacuum.

Strength of Solutions

In laboratory work various kinds of solutions are used, their difference depending on the quality of the substance dissolved. The definition of these is given below.

A saturated solution is a solution containing as much of the substance as it can possibly dissolve at a given temperature.

A concentrated solution is one which is almost but not quite a saturated one. For example, concentrated sulphuric acid contains 95 to 98 parts of acid in every 100 parts of liquid.

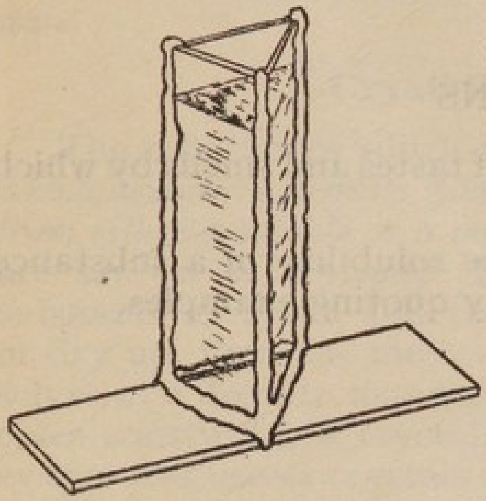
A dilute solution is one in which the quantity of the substance dissolved is very small in proportion to the solvent. For example, vinegar is a weak or dilute solution of acetic acid.

A percentage solution is a solution of a definite weight of a substance in 100 *volumes of the liquid.* For example, a 10 per cent. solution contains 10 grammes of the substance dissolved in the liquid and then diluted until the mixture measures 100 cubic centimetres.

PRACTICAL CHEMISTRY—CHAPTER 2

Colour

Expt. 7



Hold a glass prism to the light and note the primary colours—blue, yellow, and red, and the secondary colours—green, orange, indigo, and violet.

If a glass prism, such as a knife rest, cannot be obtained, one can be constructed by sealing together the edges of three miscroscope slides with sealing wax, fixing them at their base to another slide as shown in the diagram and then filling the space formed with water.

Taste

Expt. 8

Appreciate bitter, sweet, acid, and saline tastes by tasting quinine, sugar, vinegar, and salt.

Smell

Expt. 9

Distinguish aromatic, pungent, and fœtid odours, by smelling tincture of balsam, chlorine or sulphur fumes, and asafœtida.

Solubility

Expt. 10

Nature of Substance

Dissolve a quarter of a teaspoonful of salt in a test tube half filled with water, and then show the insolubility of sand, sulphur, and magnesium carbonate.

Expt. 11

Nature of Solvent

Dissolve potassium nitrate (saltpetre) in water and try to dissolve it in alcohol.

Similarly show that fat is insoluble in water yet soluble in ether.

Expt. 12

Degree of Temperature

Dissolve as much saltpetre as possible in cold water. Then decant off the saturated solution into another test tube. Heat the solution and try to dissolve more saltpetre in it.

Allow the solution to cool and watch the crystals form.

QUESTIONS

Question 1.—What are the chief tastes and smells by which substances can be identified ?

Question 2.—On what does the solubility of a substance depend ? Illustrate your answer by quoting examples.

CHAPTER 3

THE FORMS OF SOLID MATTER

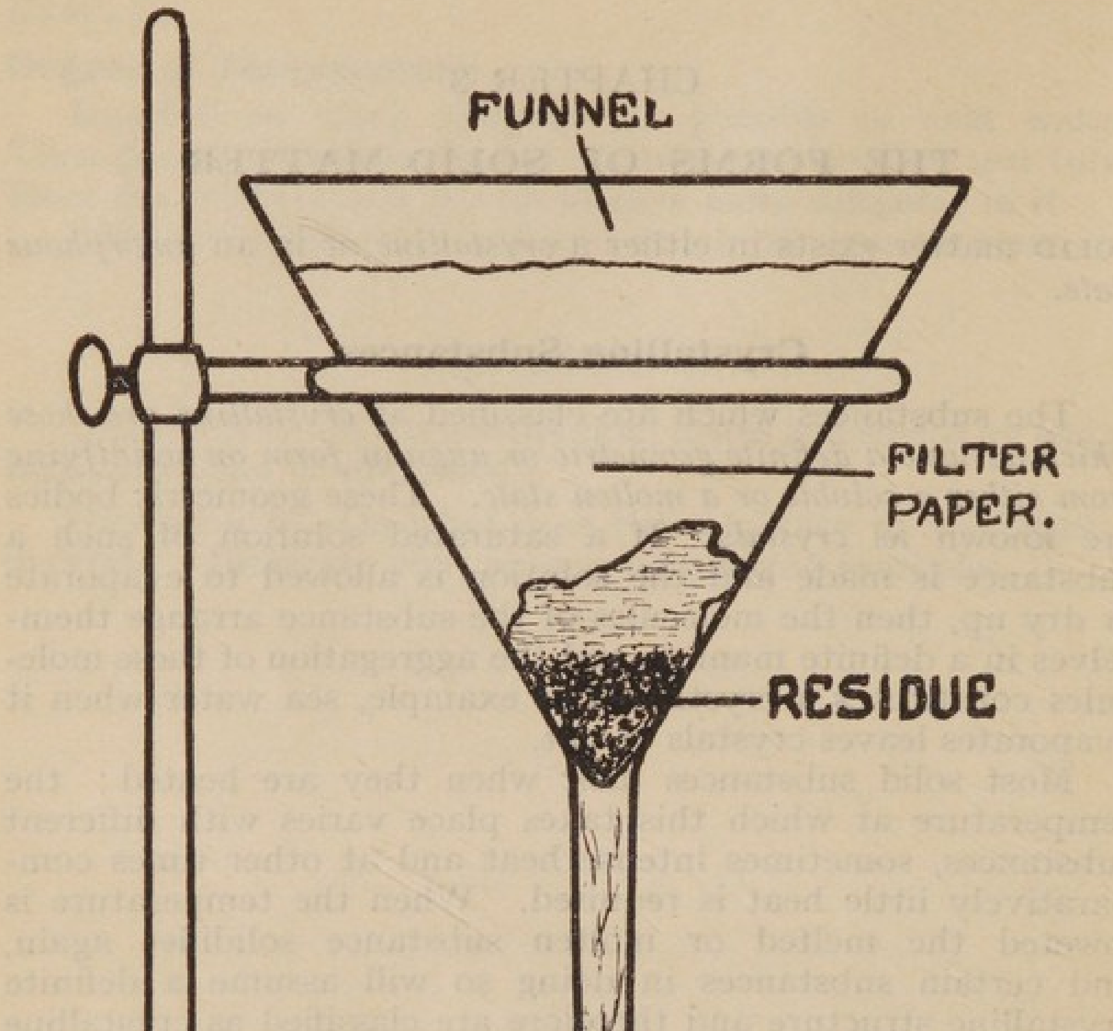
SOLID matter exists in either a *crystalline* or in an *amorphous state*.

Crystalline Substances

The substances which are classified as *crystalline are those which assume a definite geometric or angular form on solidifying from either a soluble or a molten state.* These geometric bodies are known as *crystals.* If a saturated solution of such a substance is made and the solution is allowed to evaporate or dry up, then the molecules of the substance arrange themselves in a definite manner, and the aggregation of these molecules constitutes a crystal. For example, sea water when it evaporates leaves crystals of salt.

Most solid substances melt when they are heated ; the temperature at which this takes place varies with different substances, sometimes intense heat and at other times comparatively little heat is required. When the temperature is lowered the melted or molten substance solidifies again, and certain substances in doing so will assume a definite crystalline structure and therefore are classified as crystalline substances. For example, quartz melts at a very high temperature, and when it cools it crystallizes into such forms as rock crystals and amethysts which are found in nature. The heat in the centre of the earth is sufficient to melt the quartz, and in this molten condition it is often forced up by volcanic eruption to the surface and on cooling slowly forms crystals. Crystals vary in shape from simple cubes, prisms, and needles to the most complicated formations. Salt crystallizes as cubes, potassium permanganate as prisms, and corrosive sublimate in the form of needles. They all possess smooth faces or facets, which make them sparkle in the light. Usually the crystals of a particular substance never vary in their shape, although they may in their size. Salt crystals are never found assuming any other form than cubes, potassium permanganate no other form than prisms.

The fact that certain substances possess the property of forming crystals has been made use of in the purification of these substances. The process briefly is as follows. The substance is dissolved in water and any insoluble and extraneous matter removed by *filtration,* which usually is effected by passing the liquid mixture through blotting-



FILTRATION

paper, cotton wool, linen, etc. The solid insoluble matter having been strained off, the liquid which has passed through the filter and is known as the *filtrate* is then evaporated to a sufficient concentration by applying heat, and when it is allowed to cool most of the substance will then crystallize out. These crystals are pure samples of the substance.

Water of Crystallization

Crystals, in forming, usually absorb a certain amount of water, which becomes so closely combined with the substance that it is seldom evident and is only driven off by heat. This water is known as *water of crystallization,* and crystals of different substances contain definite amounts. For example, the crystals of washing soda (sodium carbonate) contain equal quantities of soda and water, those of copper sulphate contain two parts of the solid and one of water.

Although most soluble crystalline substances contain *water of crystallization,* the crystals of salt and corrosive sublimate are composed of pure sodium chloride and pure mercuric chloride with no water in combination.

The presence of *water of crystallization* is of the utmost importance in pharmacy because certain drugs may exist in two different forms—namely, as a crystal containing *water of crystallization* or as a powder from which all the water has been driven off by heat. If a prescription contains a definite amount of such a substance and it is not stated in which form it is to be added, then there will be a considerable difference in the dose of the drug.

For example, sodium carbonate may be dispensed as crystals or as a powder. A 10-grain dose of the former would only contain 5 grains of pure sodium carbonate, the other 5 grains being water, whereas if the powder is used a full 10-grain dose of sodium carbonate would be given. It is very necessary, therefore, when a substance exists in two forms, to state definitely in the prescription which form is desired.

Amorphous Substances

It has already been stated that when a crystalline substance containing water of crystallization is heated, this water is driven off and the characteristic shape of the crystals disappears leaving a powder, *the particles of which have no definite shape or form,* and because they are without shape they are called *amorphous.*

A crystalline substance may be reduced to a powder by pounding, but however fine the powder is, its crystalline structure remains, and the powder would not therefore be amorphous.

Many substances only exist in the amorphous state, for example, powdered chalk and the oxides of certain metals such as zinc oxide or iron oxide. Powdered organic substances are commonly amorphous.

Colloids and Crystalloids

It is very necessary in connection with the study of chemistry in relation to dispensing to realize that substances can be further sub-divided into two groups depending on their capability or otherwise of passing through a *membrane*.

Certain substances in solution can pass freely through a parchment or animal membrane, they include all the crystalline salts and are known as *crystalloids ; other substances diffuse extremely slowly or not at all*, and are known as *colloids*.

The following simple experiment illustrates this property. If solutions of starch, a white colloid, and iodine, a brown crystalloid, are mixed, a deep blue colour results. If this blue mixture is placed in a parchment saucer floated in a beaker of water, the crystalline iodine will diffuse through the membrane and will stain the water brown. No starch has passed through, as the merest trace of it in the presence of iodine would be detected by a blue colour in the water. Starch is therefore a colloid.

There are two other methods of distinguishing colloids from crystalloids, namely, by examining the liquid containing them under a microscope or by holding the liquid to a strong light.

If the liquid consists of a crystalloid in solution, no particles can be seen even when viewed through a microscope of the highest magnification ; if, on the other hand, it consists of a colloid, then floating about in the liquid will be seen minute particles of the substance. When a test tube containing the former is held up in a beam of light in a dark room it looks perfectly clear, whereas the latter would appear opaque, owing to the reflection by the particles of the rays of light.

For example, glue is a colloid. In the dry form, if struck with a hammer it splinters into many sharp fragments ; when, however, a piece is placed in water it becomes soft and jelly-like, and on the addition of increasing amounts of water the jelly will become so thin that it appears almost as transparent as water. The glue has not dissolved in water, for minute particles of it can be detected when viewed under a microscope, and when the liquid is held up to a strong light it will appear opaque.

Allotropy

Many elements exist both in a crystalline and in an amorphous state. This property of occurring in the two distinct forms goes by the name of *allotropy*, and the forms the element assumes are said to be *allotropic modifications* of the element. For example, carbon may exist as a diamond crystal or as charcoal powder. They are both pure forms of carbon and are known as Allotropic Modifications of Carbon. Similarly, sulphur may occur as a yellow brittle substance and as an amber-coloured elastic substance, yet they are both forms of pure sulphur.

The element phosphorus exists in two distinct forms, namely, red and white phosphorus, which are not only entirely dissimilar in appearance but also different in their chemical properties. The two forms are allotropic modifications of pure phosphorus.

PRACTICAL CHEMISTRY—CHAPTER 3

Crystalline Substances

EXPT. 13

Form of Crystals.—Examine specimens of crystals of potassium permanganate, copper sulphate, and common salt.

EXPT. 14

Solubility. Repeat experiment 12.

Water of Crystallization

EXPT. 15

Heat some sodium carbonate or copper sulphate crystals in a test tube and note the change. The water of crystallization is driven off, and a non-crystalline powder results.

Now heat a little pure salt and note no change, because its crystals contain no water of crystallization.

Amorphous Powder

EXPT. 16

Take some copper sulphate crystals and pound them into the finest powder in a mortar. Examine with a magnifying glass and note that the crystalline form is still maintained.

Now heat the crystalline powder, driving off all water of crystallization, and again examine with a lens—note the absence of all crystalline form. The copper sulphate is now in an amorphous state.

Examine flowers of sulphur and powdered charcoal with a lens, both will be found to be amorphous.

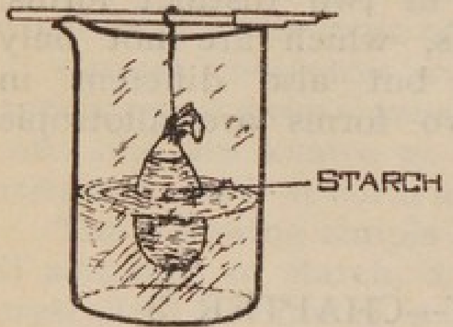
Colloids and Crystalloids

Expt. 17

Make up a thin starch solution by boiling a pinch of starch in a test tube of water. Hold the test tube up to the light and note opacity of the solution.

Expt. 18

Place the starch solution in a small sac made from sausage skin or meat membrane and suspend in water contained in a beaker as shown in diagram, and note that none of the starch solution passes through because starch is a colloid. The absence of starch in the water outside the sac may be shown by its failure to give a blue colour when a solution of iodine is added.



Expt. 19

Repeat this experiment, using salt solution, and from time to time test samples of the water in the beaker for salinity by adding some silver nitrate solution, by looking for a white precipitate which indicates the presence of salt. The salt passes through the membrane because it is a crystalloid.

QUESTIONS

Question 1.—What is the meaning of the terms "crystalline" and "amorphous" as applied to substances ?

Name the characteristics and give examples of each.

Question 2.—What do you understand by " water of crystallization " and of what importance is it from the point of view of dispensing ?

Question 3.—What are the characteristics of crystalloids and colloids ? Give examples of each.

Describe a method of separating them when mixed together.

CHAPTER 4

LIQUID SUBSTANCES

Cohesion

The term *cohesion* has been defined and mention has been made of it in relation to liquids. It has also been noted that the molecules of a liquid possess sufficient cohesion to keep them in contact, yet allowing them to slip or glide over one another. In other words, all liquids flow, but the ease with which they flow varies with the nature of the fluid.

Viscosity

There is a great difference between thick lubricating oil and petrol. The reason depends upon the greater cohesion of the molecules of oil compared with those of petrol; they stick together more, and fluids of this description are called *viscous* or thick. The term *viscosity* is used to denote this characteristic of liquids. Petrol is not viscous but belongs to the thin or mobile class of liquids, which also includes spirits and water. The mobility or viscosity of a liquid therefore depends on the degree of cohesion of its molecules. On account of its greater cohesion, oil can be poured in a much finer stream than water, which soon breaks up into drops.

Meniscus

The characteristic property of a liquid, namely, of taking the shape of the vessel into which it is poured, has also been stated. To continue further, all liquids maintain an even or flat upper surface so long as the surface is at rest and is free from contact with the vessel. The surface must be sufficiently large, for, if liquids are confined in narrow tubes, the surface, or, as it is sometimes called, the *meniscus*, no longer remains level, but may be concave, with the curved surface directed downwards, or convex, with it directed upwards. Whichever it is, depends on whether the liquid wets the surface of the tube or not.

If it does not wet the surface, as in the case of mercury, the meniscus is convex, whereas if it does, like water, then the molecules on the tube pull those on the surface of the liquid, with the result that it is pulled up at the edge, and the surface when viewed from the side appears concave.

Capillary Attraction

The force or power which draws the fluid up is known as *surface tension* or *capillary attraction.* If the tube has a very narrow bore the fluid may be drawn up a considerable distance, and this is made use of in practice in taking samples of blood for examination.

Another example of surface tension is seen in the case of bubbles on the surface of a liquid. If the liquid is water the bubbles burst almost immediately after forming, the reason being that the surface tension, or pull on the molecule, is greater than their cohesion. If some soap solution is added to the water the surface tension is now less than the cohesion of the molecules and the bubbles can be separated from the surface of the fluid without bursting.

The practical value of this is sometimes seen when frothing of a mixture is troublesome. A little alcohol added to the fluid reduces its viscosity and bubbles cease to form.

Osmosis

It has been stated in Chapter 3 that one of the main differences between colloids and crystalloids is the property the latter substance possesses of passing or diffusing through an animal or parchment membrane. The scientific term for this kind of diffusion is *osmosis.* It is by osmosis that some of the fluids of the human body pass from one tissue to another.

If two fluids capable of being mixed are placed in a glass vessel, the heavier one first and the lighter run gently on the top, the liquids will diffuse one into the other until the solution is a uniform mixture. If a membrane such as is used for sausage skins had been interposed between the original fluids, diffusion would still have continued, although more slowly, with eventually the same result—a uniform mixture above and below the membrane. This would be an example of diffusion by osmosis. So long as the solutions are passing one into the other there is a current set up, and this current or pressure is called *osmotic pressure.* It ceases when the mixture is complete. Take, for example, a solution of iodine, which is brown, and a salt solution, which is colourless ; if they are separated by a parchment diaphragm the iodine being crystalline will pass into the salt solution, and the salt being also crystalline will pass into the iodine. The currents, or osmotic pressure, will be evident by the " streaming " of the brown iodine into the salt solution. Eventually both solutions will be the same depth of brown and the same degree of salinity, showing that osmosis has taken place. The solutions above and below the diaphragm will be similar in every respect.

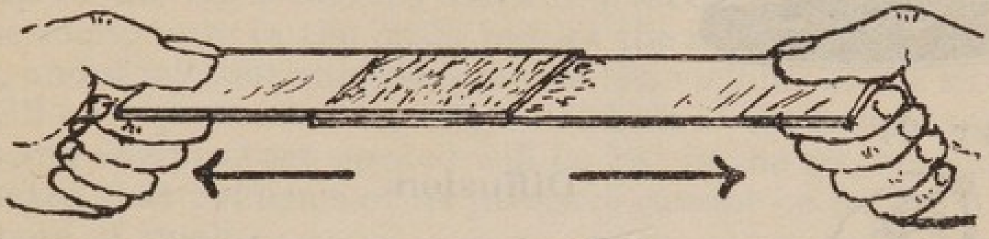
PRACTICAL CHEMISTRY—CHAPTER 4

EXPT. 20

Cohesion

Press a drop of thick oil between a pair of microscope slides, excluding air bubbles, and repeat, using water in place of oil.

Note, that when attempting to separate the slides by drawing them apart in a horizontal direction, those held together with oil are much harder to separate than those moistened with water.



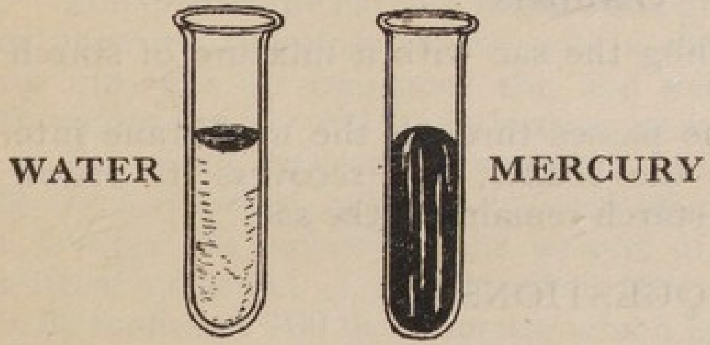
Viscosity

Suspend an empty tin, in the bottom of which a small hole has been punched, and fill it with water. Note that the water escapes rapidly in the form of drops from the tin.

Empty the tin and repeat the experiment, using oil. The oil will very slowly flow through the aperture and will escape as a thin unbroken stream.

EXPT. 21

Meniscus



Compare the meniscus in test tubes containing mercury and water.

Note, that in the former the meniscus is convex and that the tube remains dry, while in the latter the meniscus is concave and the tube wet.

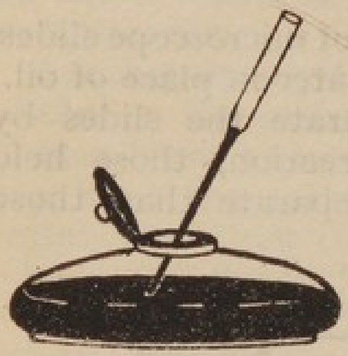
EXPT. 22

Surface Tension

Shake up a little soap with some water in a test tube; decant off ¼-inch of the soapy water into another test tube. Dilute this with distilled water until, on shaking, a ¼-inch of froth remains. Now add some methylated spirit. Note the disappearance of the froth, which cannot be reformed on shaking if sufficient methylated spirit is added.

EXPT. 23

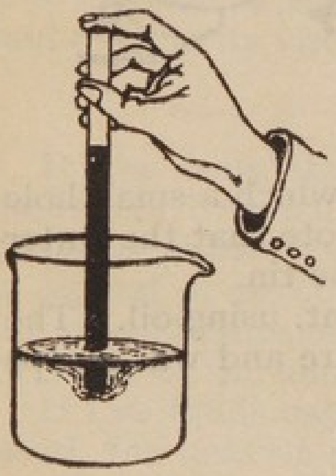
Capillary Attraction



Demonstrate this by using a dropper eye glass or a fountain-pen filler and some ink.

Dip the point of the tube into the ink and note that the fluid runs up the narrow point of the tube.

EXPT. 24



Diffusion

Gently run into a beaker, containing some salt solution, a small quantity of permanganate solution.

Leave for several hours, and note how the fluids gradually diffuse one into the other, until the solution is of an even colour throughout.

EXPT. 25

Osmosis

Repeat Expt. 18, filling the sac with a mixture of starch and iodine solution.

The crystalloid-iodine passes through the membrane into the water contained in the beaker, and recovers its brown colour, while the colloid starch remains in the sac.

QUESTIONS

QUESTION 1.—Explain the terms "viscosity," "meniscus" and "capillary attraction," as applied to liquids.

Describe an experiment illustrating each.

QUESTION 2.—What is the difference between diffusion of liquids and osmosis?

Amplify your answer by describing two experiments which demonstrate this difference.

CHAPTER 5

GASES AND THEIR PROPERTIES

The Effects of Pressure

THE molecules of gas possess no cohesion; in fact, the very opposite state of affairs exists: they repel one another, and do so to such an extent that if not prevented by some force, they will to all intents and purposes be lost in the air. Leave the tap in the bath running for a few seconds and the water will be found only in the bath, but let the gas escape from the stove, and it will not only be found in the vicinity of the stove, but will have permeated rapidly the whole of the house, and if doors and windows are opened will be lost in the air, becoming so diluted that evidence of its presence cannot be detected by the sense of smell.

Gas in gasworks does not escape, because it is enclosed in huge cylinders, nevertheless, it is constantly pressing on the walls trying to get free. Gases, therefore, exert pressure and the closer their molecules the greater the pressure will be.

The air is a mixture of gases and obeys all the laws of gases. Its individual molecules are kept moderately closely packed and are prevented from straying widely by the pressure of weight of the atmosphere. The air, in other words, is compressed by its own weight in the same way as sand by the pressure of millions of tons of sand above may be compressed into sandstone.

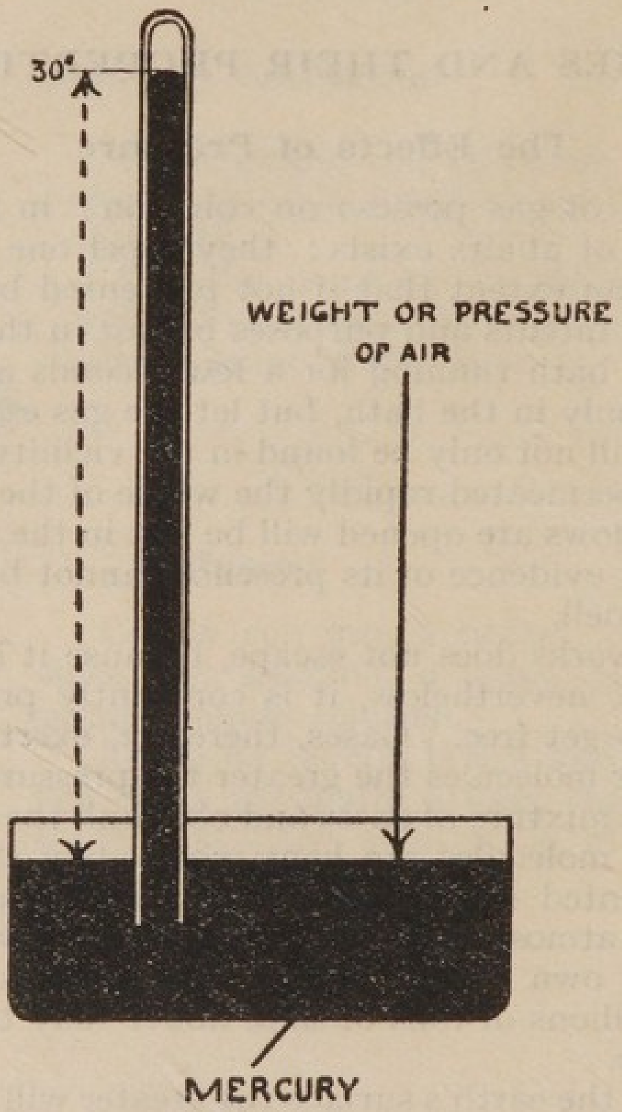
The nearer the earth's surface the greater will be the weight or pressure, therefore the air will be densest here; the further up into the air one goes the less will be the pressure and therefore the air will be thinner or more " rarefied."

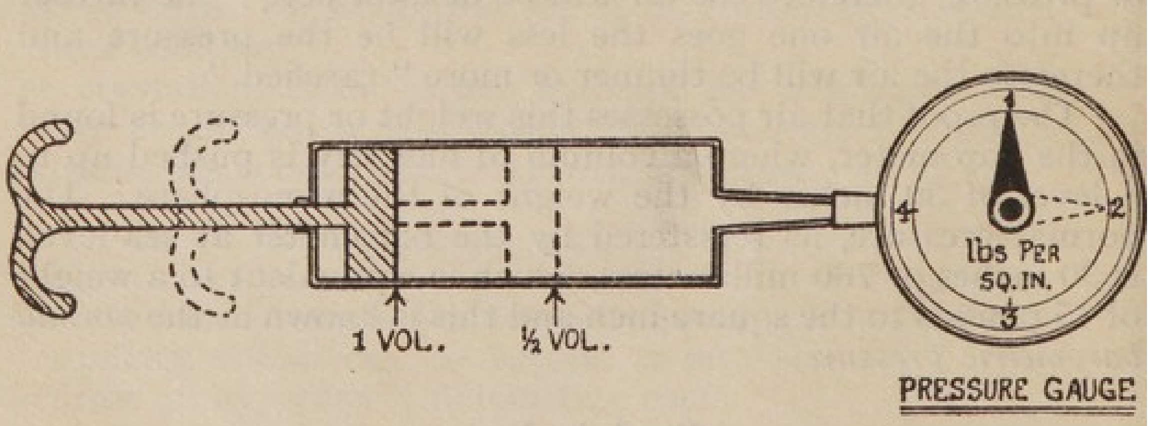
The proof that air possesses this weight or pressure is found in the barometer, where a column of mercury is pushed up to a level of 30 inches by the weight of the atmosphere. The normal pressure, as registered by the barometer at sea-level, is 30 inches or 760 millimetres, which is equivalent to a weight of 15 pounds to the square inch and this is known as the *normal barometric pressure*.

Boyle's Law

Now this pressure has a very important effect on the behaviour of gases, so important that one, Boyle by name, made a definite statement about it, which statement is now known as " Boyle's Law." In a few words it amounts to this:

If the pressure is doubled the volume of the gas is halved, if the pressure is halved the volume of the gas is doubled, so long as the





gas remains at the same temperature. For example, supposing some air is drawn up into an air syringe, to the nozzle of which is fixed a pressure-gauge which registers the pressure exerted by the volume of the air in the syringe. If the piston is forced down so that this volume of air is reduced to half then the gauge will register double the pressure, and vice versa if the pressure is doubled the volume of air will be found compressed to half. Now if the piston is withdrawn so that the volume of air is once again doubled, then the pressure as indicated by the gauge will have dropped to half.

With increasing pressure the molecules of a gas are packed closer and closer together, and if at the same time it is cooled down to a certain temperature which varies with the gas, then the gas changes its state and becomes a liquid.

Effects of Heat

The next important factor affecting a gas is *heat*. What effect has it ? It is well known that when most substances are heated they expand; iron expands and that is why rails are laid with a slight gap between them so that on a hot summer's day they can expand without buckling; water expands, a kettle full of cold water when heated will overflow, due to expansion.

Charles's Law

Gases expand even more than liquids or solids and they do so in a regular manner according to the amount of heat applied. A scientist, Charles, discovered this fact and set it out in what is now known as " Charles's Law," namely, that the *Volume of all gases increases or diminishes by the same amount for every degree rise or fall in temperature so long as the pressure remains the same.*

Diffusion

Gases *diffuse* one into another just as liquids have been described as doing, only they do so with greater rapidity.

No two gases have the same weight. Air, although a mixture of gases, is taken as the standard gas with which the other gases are compared, being either heavier or lighter.

The heavier gases diffuse slowly, the lighter rapidly ; the reason being that the molecules of the latter are moving about more actively and exert a greater repelling force on each other than those of the heavier gases, the behaviour of which may be compared with that of the more viscous fluids. To illustrate these points, take as an example the two gases, hydrogen, which is a colourless light gas, and nitrogen peroxide, which is a reddish-brown heavy gas.

If two glass vessels are filled with these gases, the one with hydrogen, and the other with nitrogen peroxide, and the hydrogen vessel placed mouth downwards over that containing the nitrogen peroxide, this brown heavy gas will be seen in spite of its weight to pass up into the clear hydrogen gas, which will pass down, and eventually both vessels will contain a uniform mixture of the two gases. Were the heavy gas put into the inverted vessel then the mixing would continue more rapidly, for in addition to the action of diffusion the heavy gas would sink and the light gas would rise.

The fact that gases have different rates of diffusion has been made use of in order to separate a mixture of them. This is done by placing the mixture in a porous pot, when the lighter gas will diffuse rapidly through the pot, leaving the heavier one inside.

Solubility of Gases

Most gases are soluble in water or alcohol and their *solubility* depends on four factors :

(1) The nature of the liquid.
(2) The nature of the gas.
(3) The temperature of the liquid.
(4) The pressure.

The Nature of the Liquid.—Some liquids dissolve more gas than others. For example, alcohol can dissolve roughly 3½ times the quantity of carbonic acid that water can.

The Nature of the Gas.—Oxygen dissolves sparingly in water, whereas ammonia is very soluble indeed, something like 28,700 times more so than oxygen.

The Temperature of the Liquid.—The hotter the liquid the less gas it will hold in solution. Boil water and all gases will be expelled.

The Pressure.—Only a definite volume of gas can be dissolved by a liquid, so that by forcing more gas in by pressure there is no alteration in the actual volume, but although this remains the same the *weight* of the gas is increased in proportion to the pressure. To put it more simply, supposing one cupful of gas is the most at ordinary times a pail of water will dissolve, if another cupful is forced in under pressure, the gas in the water will only occupy the space of one cupful. There are two cupfuls of gas compressed into one, so that though the volume of gas would be the same the weight of the gas dissolved would be double.

PRACTICAL CHEMISTRY—CHAPTER 5

Gases

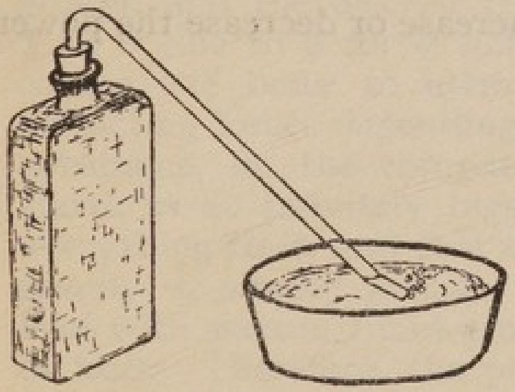
Expt. 26

Pressure

Demonstrate this in a rough way by lightly filling a rubber balloon with air—noting the pressure on the rubber. Then grasp half the balloon, reducing its size by half, note that the pressure in the rubber envelope is doubled.

Expt. 27

Heat



Fit a tight cork into an ordinary medicine bottle. Through the cork push one end of a glass tube bent so that the other end can be immersed in a basin of water.

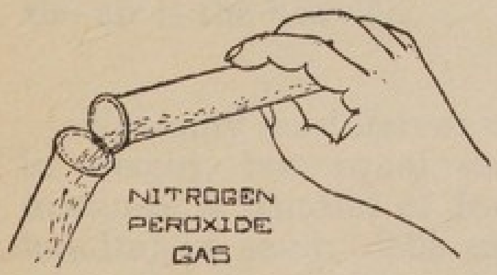
Warm the bottle with the hands, and note the bubbles of air passing through the water due to expansion of the air.

Cease warming, and note the water is drawn up by the contraction of the air in the bottle.

Expt. 28

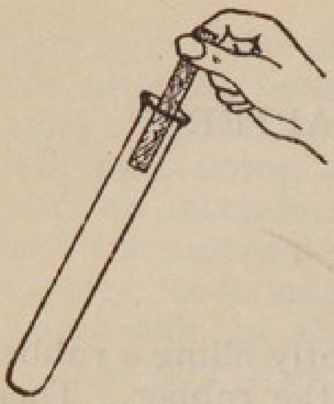
Diffusion

Add a few drops of strong nitric acid to a few pieces of copper wire in a test tube (be careful when pouring—nitric acid is dangerous). A brown gas comes off, which is nitrogen peroxide. Collect some of the gas by inclining the tube containing it over an empty tube.



Now invert an empty tube over the tube containing the gas and note the rate of diffusion. As the gas is heavy, diffusion will be slow.

Repeat the experiment, only inverting the tube containing the gas over the empty tube. Diffusion is much more rapid because the weight of the gas assists in the mixing.



EXPT. 29

Solubility

Repeat Expt. 3. Ammonia gas is very soluble in water, but if the solution is heated the ammonia will be driven off and its presence can be demonstrated by the use of some red litmus paper.

QUESTIONS

QUESTION 1.—What effects have pressure and temperature on the behaviour of gases ?

QUESTION 2.—Why do gases diffuse, and on what does the rate of their diffusion depend ?

QUESTION 3.—What factors increase or decrease the power of fluids to dissolve gases ?

CHAPTER 6

WEIGHT—DENSITY—SPECIFIC GRAVITY

A THOROUGH knowledge of weights and measures is essential to the dispenser, for the slightest mistake on his part when dealing with poisonous drugs may mean the death of the patient. It is not intended to describe here the various kinds of weights and measures, these will be found in any book on dispensing, but to explain in general terms what the expressions Weight, Density, and Specific Gravity mean.

Weight

When a body is dropped from a height it will fall to the ground and the heavier the body is the quicker it will fall. This is due to a force which exists in all matter and which causes one body to attract another, the strength of the attracting forces depending, if both are composed of the same substance, on the comparative size of the bodies. As the Earth is so infinitely bigger than anything else it attracts everything to it, and this attracting force is called the *force of gravity* or simply *gravity*. Gravity will attract a pound of lead with sixteen times greater strength than it will an ounce of lead. Therefore, the pound is sixteen times heavier than the ounce. If one holds the pound in one hand and the ounce in the other, this pull or attraction will be felt and goes by the name of *weight*.

Certain gases, for instance, hydrogen gas, will not fall to the ground, but will float higher and higher into the air. Gravity is nevertheless attracting the hydrogen, but just as a cork in the water will float to the surface and defy gravity, so does hydrogen float on the surface of the air, because in the case of the cork the water is the heavier and in the latter case the air is the heavier.

Density

Not only do different volumes of the same substance differ in weight, but equal volumes of different substances; for instance, a handful of feathers will feel much lighter than a handful of sand. The explanation of this becomes evident when the composition of matter is again examined. Matter is composed of molecules, and the molecules of atoms, and the closer the atoms are packed together the greater is the *density* of the substance, and as all dense matter is heavy therefore the weight of the substance will also be greater. There is in this way a relationship between *density* and *weight*, although they refer to different things. Density refers to the composition of the matter, and weight to the force the Earth has of attracting matter. The same word "dense" is applied to a fog and to a population; in the former the particles of moisture that compose the fog are so closely packed that visibility is great reduced (restricted to only a few yards), and in the latter density implies that people are congregated and packed together in dwellings in towns.

Specific Gravity

It is often necessary when making up mixtures to know the comparative weights of the various substances used, a *standard* substance is therefore selected with which all substances can be compared. The standard for solids and liquids is usually the weight of pure water at 4 degrees Centigrade, because at this temperature water is at its heaviest; the standard for gases is generally air, or hydrogen because this gas is the lightest substance known.

The question to be answered is, how does the weight of any substance compare with the weight of a similar volume of water, or if the substances are gases, with a similar volume of air or hydrogen? The answer in the case of the solids and liquids is based on the Law of Archimedes, who many hundreds of years ago discovered that when a body is immersed in water, the body loses weight and the weight lost is equal to the weight of an equal volume of water. That a body loses weight in water is evident to any one who tries to raise a large stone out of water. The stone can be lifted with comparative ease so long as it is immersed in the water, but when it appears above the surface its weight seems to increase until once clear of the water it may be so heavy that it cannot be carried. The reason for this is that water exerts a force in an upward direction and assists the person in raising the stone.

When the body is lighter than water, the upward force of water is greater than the downward force of gravity acting on the body, and so the body floats to the surface. In the case of the stone, the amount the stone is lighter in water is the weight of an equal volume of water. For instance, suppose the weight of the body in air is 14 ounces, and, immersed in water, it is 12 ounces, then the weight of an equal volume of water would be 2 oz., and the *specific gravity*, that is the comparative weight of the body with water, would be seven times more or it would have a specific gravity of 7.

A simple experiment to find the specific gravity of, for example, lead is as follows. The lead is suspended by a fine silk thread from one arm of the balance, weighed and its weight noted, a beaker of water is now placed on a platform so that when the weight is immersed it hangs suspended in the water. It will be found that the weights will require reducing, and the amount by which they are reduced is the weight of an equal volume of water. When one is dealing with substances that float, then weights have to be added to sink them and the experiment becomes somewhat more complicated in consequence. To determine the specific gravity of liquids, *hydrometers* are used. They are merely weighted floats, and the deeper they dip the lower the specific gravity of the fluid, the higher they float the greater the specific gravity. There are special hydrometers for different kinds of fluid, for example, a lactometer for finding the specific gravity of milk and a urinometer for urine.

PRACTICAL CHEMISTRY—CHAPTER 6

Density

Expt. 30

Density of Liquids

Fill an empty ether bottle with a saturated solution of magnesium sulphate and compare its weight with that of a full bottle of ether.

The solution of magnesium sulphate, although equal in volume to the ether, is much heavier.

Expt. 31

Density of Solids

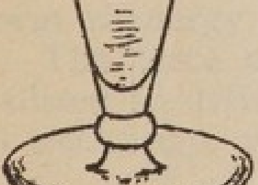
This can be demonstrated by comparing two substances of equal volume, such as cork and lead, aluminium and silver.

Expt. 32

Density of Gases

This was shown in Expt. 28, where the nitrogen peroxide, by its greater weight, could be poured like a liquid from one tube to another.

Specfic Gravity



Expt. 33

Specific Gravity of Liquids

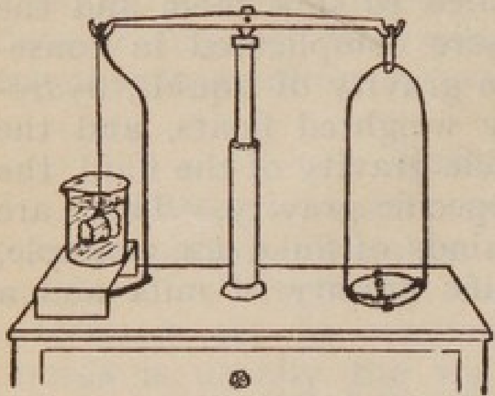
Float a lactometer or urinometer first in water, then in magnesium sulphate solution, in a 20-oz. measure (taking care that it does not touch the sides).

Note that the lactometer will float at a higher level in the latter, because its specific gravity is the greater.

Expt. 34

Specific Gravity of Solids

Weigh a piece of lead or copper in the air. Weigh again (in water) whilst suspended from the beam (see diagram).



Subtract its weight in water from its weight in air. This gives the weight of an equal volume of water.

The question is how much greater is the dry weight of the solid than the weight of an equal volume of water, which is taken as the standard with which all weights are compared.

The dry weight of the solid therefore is divided by the weight of the water.

Example :

Dry weight .. .. .. = 30 oz.
Weight in water .. .. = 27 oz.

Loss of weight or the weight of an equal volume of water = 3 oz.

$$\frac{\text{Dry weight}}{\text{Weight of the water}} = \frac{30}{3} = 10$$

The solid was 10 times heavier than water, therefore its specific gravity is 10.

QUESTIONS

Question 1.—In comparing two different bodies of equal size, one is found to be heavier than the other. What is the scientific reason for this ?

Question 2.—The specific gravity of a substance is 10. What do you understand by this ? Describe a simple experiment to determine the specific gravity of a liquid and of a solid.

CHAPTER 7

THE EFFECTS OF PHYSICAL FORCES

Heat—Physical Force

IT is necessary to explain in the first place what is meant by the term *physical force* before its effects are discussed. The word *physical* denotes something connected with matter and *force* is a power acting on a body, so as to make that body move if it is at rest, or if moving, to bring it either to a standstill or to deflect it from the line along which it travels.

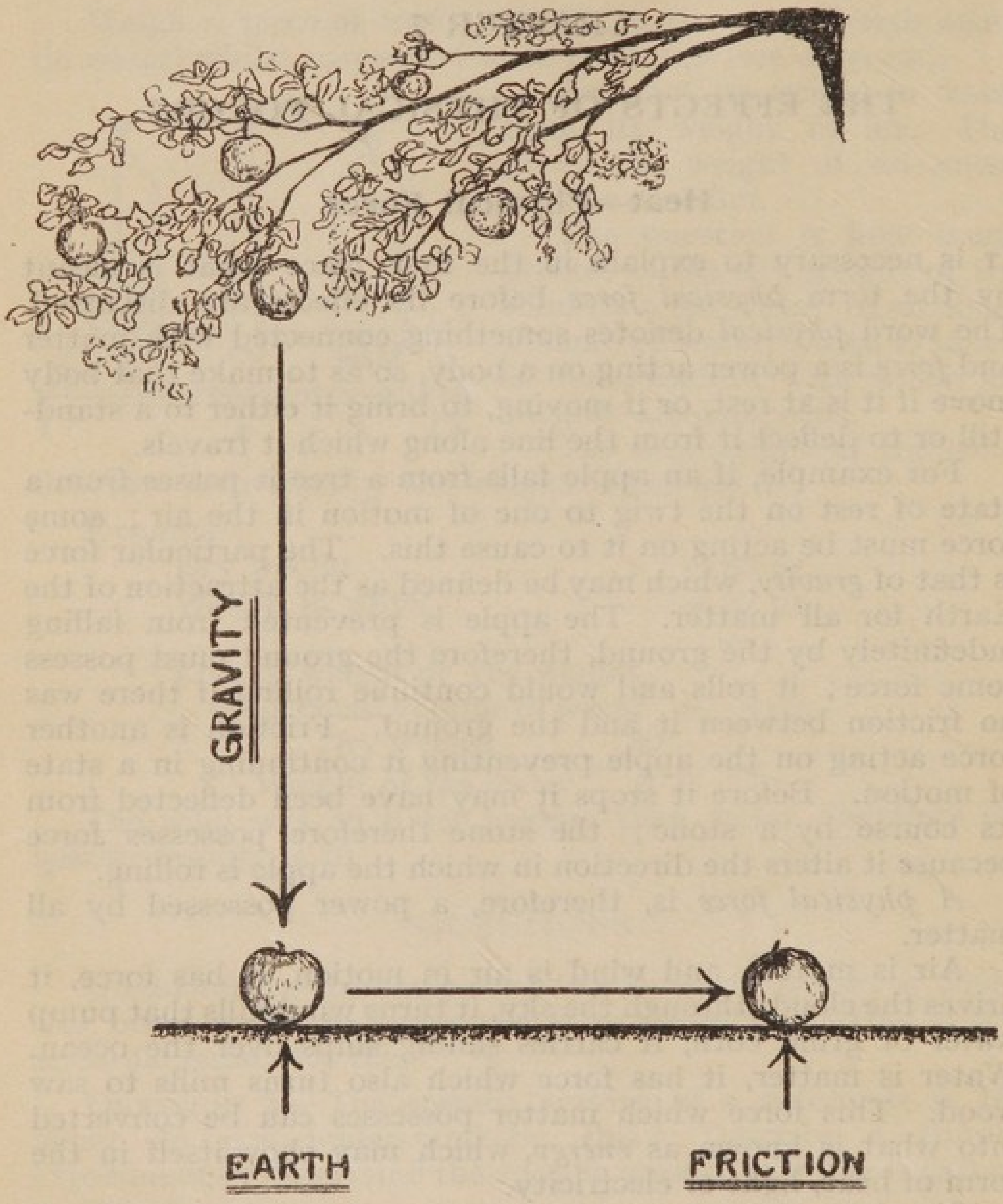
For example, if an apple falls from a tree it passes from a state of rest on the twig to one of motion in the air; some force must be acting on it to cause this. The particular force is that of *gravity*, which may be defined as the attraction of the Earth for all matter. The apple is prevented from falling indefinitely by the ground, therefore the ground must possess some force; it rolls and would continue rolling if there was no friction between it and the ground. Friction is another force acting on the apple preventing it continuing in a state of motion. Before it stops it may have been deflected from its course by a stone; the stone therefore possesses force because it alters the direction in which the apple is rolling.

A physical force is, therefore, a power possessed by all matter.

Air is matter, and wind is air in motion, it has force, it drives the clouds through the sky, it turns windmills that pump water or grind corn, it carries sailing ships over the ocean. Water is matter, it has force which also turns mills to saw wood. This force which matter possesses can be converted into what is known as *energy*, which may show itself in the form of heat, light or electricity.

If a brass button is rubbed vigorously on a cloth it will become unpleasantly hot. The force exerted has been converted into energy or heat; if the action were continued long enough the button might become so hot that the cloth would burn. Should this do so in a dark room the flame as well as giving off heat would also give light, which is another form of energy.

The force or power possessed by running water can be utilized to drive dynamos which generate energy in the form of electricity, and this electricity can be converted back again into force to drive machinery, trains, and tramcars, or it may show itself as electric heat or electric light.



Heat

It is now necessary to deal with the forms of energy more fully, commencing with *heat*.

Nobody knows exactly what heat is. Its presence is only appreciated by its effects. A thing cannot be said to be hot unless it feels so to the touch or unless by experience it is realized that all things that burn and give off smoke are hot.

Degree of Heat—The Thermometer

There are *degrees* of heat, the sun's rays are warm, boiling water is hot, and molten iron is extremely hot. To depend on one's sense of touch to tell how much the one is hotter than the other would be a very painful and inaccurate method, therefore advantage is taken of the fact that most substances when heated expand in a more or less regular manner, and by the amount of expansion one is able to tell the degree of heat.

There are two substances commonly used for this purpose, namely, mercury for high temperatures because it has a high boiling point, and alcohol for low temperatures because it has a low freezing point.

Mercury is in every way a very suitable substance, because it can be got in a pure state, expands uniformly and quickly being sensitive to alterations in temperature, and remains in the fluid state through a wide range of temperature.

The instrument which contains the fluid is called a *thermometer*. There are various kinds of thermometers, but they are all similar in the following respects, namely, they consist of a bulb which holds the fluid, either mercury or alcohol, prolonged into a narrow hollow tube, the stem, on which is marked a scale. The stem is sealed at the top and all air is removed from the inside so that the mercury is contained in a vacuum.

Centigrade and Fahrenheit Scales

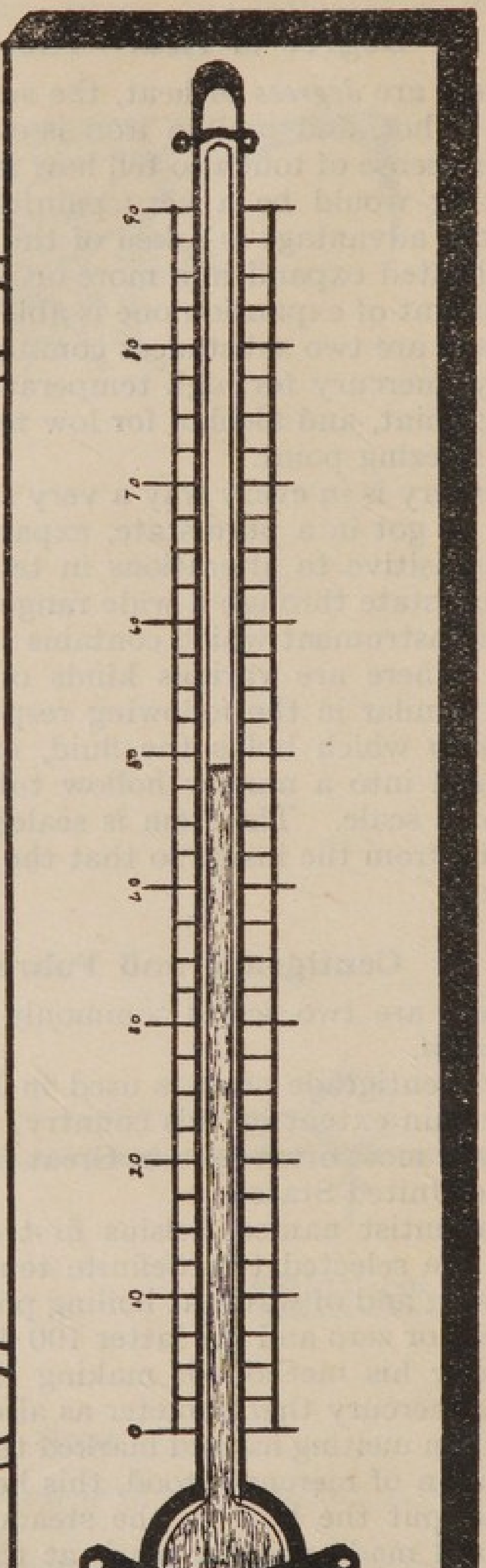
There are two scales commonly used, *Centigrade* and the *Fahrenheit.*

The Centigrade scale is used in France and Germany and to a certain extent in this country, but the Fahrenheit scale is the one most often used in Great Britain and also in Canada and the United States.

A scientist named Celsius first described the Centigrade scale. He selected two definite temperatures, that of water at freezing and of water at boiling point, and called the former 0 degrees or zero and the latter 100 degrees.

Briefly his method of making the scale was as follows: Using a mercury thermometer as already described, he buried the bulb in melting ice and marked the stem at the point where the column of mercury stood, this he called zero or 0 degrees. Then he put the bulb in the steam coming off pure boiling water and made another mark at the level the mercury rose to, this was the boiling point of 100 degrees. He then divided the interval between these two points into 100 equal divisions ; in this way a scale was produced showing temperatures between the freezing and boiling points of water.

MERCURY THERMOMETER



SPIRIT THERMOMETER

Fahrenheit took as his two fixed points: (1) the lowest temperature recorded at the time, (2) that of blood heat. He called (1) zero, and (2) 100 degrees, and divided up the interval into 100 parts. He found on his scale that the temperature of water at freezing point was 32 degrees and at boiling point 212 degrees. So he had to continue his scale from 100 up to this figure.

Quantity of Heat—Specific Heat

Substances differ in the amount of heat they can hold or store up. It requires a far longer exposure to a Bunsen flame to raise the temperature of water from 0 to 100 than to raise the temperature of an equal bulk of iron the same number of degrees. Water has absorbed more *units of heat* than iron to reach the same temperature. It is, therefore, said that the *specific heat* of water is greater than that of iron.

The specific heat of water is taken as one and that of other substances is compared with it.

Unit of Heat—The Calorie

The unit of heat is called a *calorie*. *A calorie is the amount of heat required to raise one gramme of water one degree Centigrade.* To raise a kilogramme or 1,000 grammes of water one degree Centigrade would take 1,000 calories or one kilo-calorie.

The calorie is the unit of energy as well as the unit of heat, for energy and heat are the same thing.

Latent Heat

There is another kind of heat called *latent heat*. This is the heat which is absorbed or given off when a substance changes its state either from a solid to a liquid or a liquid to a gas or vice versa.

When water boils it is converted into steam, which is a gas; it has therefore altered its physical state. So long as steam is being generated the temperature of the water will remain the same in spite of a Bunsen flame, the temperature of which is infinitely greater, playing on it continuously. All the extra units or calories of heat given off by the flame are going to convert the water into steam.

These calories compose the latent heat of steam. It requires one calorie to raise one gramme of water one degree Centigrade, but it requires 536 calories to convert this amount of water at 100 degrees to steam at 100 degrees, and when the steam condenses again the 536 calories of latent heat are set free.

In practice latent heat is of great importance. For example, in the evaporation of every gramme of sweat from the body, 536 calories are got rid of and so the heat of the body is dissipated ; also in steam disinfection of clothing when the steam condenses, 536 calories are liberated to assist in the disinfecting process.

Effects of Heat

Now what are the effects of heat on matter ? It has been shown that most substances expand. The iron rims of cart wheels are put on the wooden frame when very hot, as they expand while hot and contract on to the frame when they cool, thus binding themselves on tightly. Solids and liquids, however, do not all expand at the same rate. For every degree rise in temperature, iron expands more than glass and glass more than wood. This is unlike the behaviour of gases, which, we saw, all expand at the same rate. As more and more heat is applied to a solid it is first converted to a liquid and then to a gas, that is, if sufficient heat can be generated.

Transference of Heat

There are three ways in which heat can pass or be transferred from one object to another, namely, by *conduction, convection* and *radiation*.

Conduction

By conduction the heat is transferred from one molecule to another or from one substance to another, the molecule or substances being in contact.

If one end of a copper rod is placed in the fire, in a very short space of time the other end becomes so hot that it cannot be held by the hand. The heat from the fire has been transferred from molecule to molecule of the copper until the whole rod becomes extremely hot. Had the rod been of wood the end in the fire might burn to a cinder without there being any appreciable transference of heat to the other end.

The copper is said to have greater *conductivity* than wood. Substances can, therefore, be classed according to whether they are good or bad *conductors of heat.* Good conductors allow the heat to escape, bad conductors prevent the loss of heat.

For instance, hot water pipes may be made of iron or copper, both good conductors. Heat is transferred from the molecules of water in contact with the pipes which allow the heat to escape and so warm the air. Were the pipes made of thick wood or were they covered with a thick layer of flannel, both examples of bad conductors, no heat would come through to the surface, for heat is prevented from passing from molecule to molecule, owing to the particular nature of the pipes or covering, and therefore there would be no heating of the room.

Dry cotton wool, sawdust, and asbestos are all bad conductors. They are used to line hot boxes or ice boxes, in the one case to prevent loss of heat from the hot food inside, and in the other case to stop the heat from outside passing in and melting the ice. To prevent a hot pan or teapot from marking a polished table an asbestos mat is sometimes placed under them, and in this way transference of heat is stopped.

Convection

Molecules of air in contact with any hot material, for instance, burning coal in the grate or the iron surface of a stove, will be heated by *conduction*, the air will expand and become lighter, and if there is a flue or a chimney will pass rapidly up causing a current of hot air which will warm other objects, by coming in contact with them. The method of distributing heat by currents is called *convection*.

It is in this way that air in a room is heated, and it is also the way liquids are heated. When a kettle of water is put on the fire the water nearest the flames is raised to a higher temperature than the rest of the water, expands and becomes lighter, "streaming" to the surface as a current while the colder water on the surface runs down to take its place, where it in its turn becomes hot and rises. These convection currents continue until the water is uniformly raised to 100° C., when it boils.

Radiation

The third method of heat transference is by *radiation*. This is the opposite of conduction and convection, for there is no contact between the molecules or the substances and the intervening air is not affected so that no convection currents can be formed.

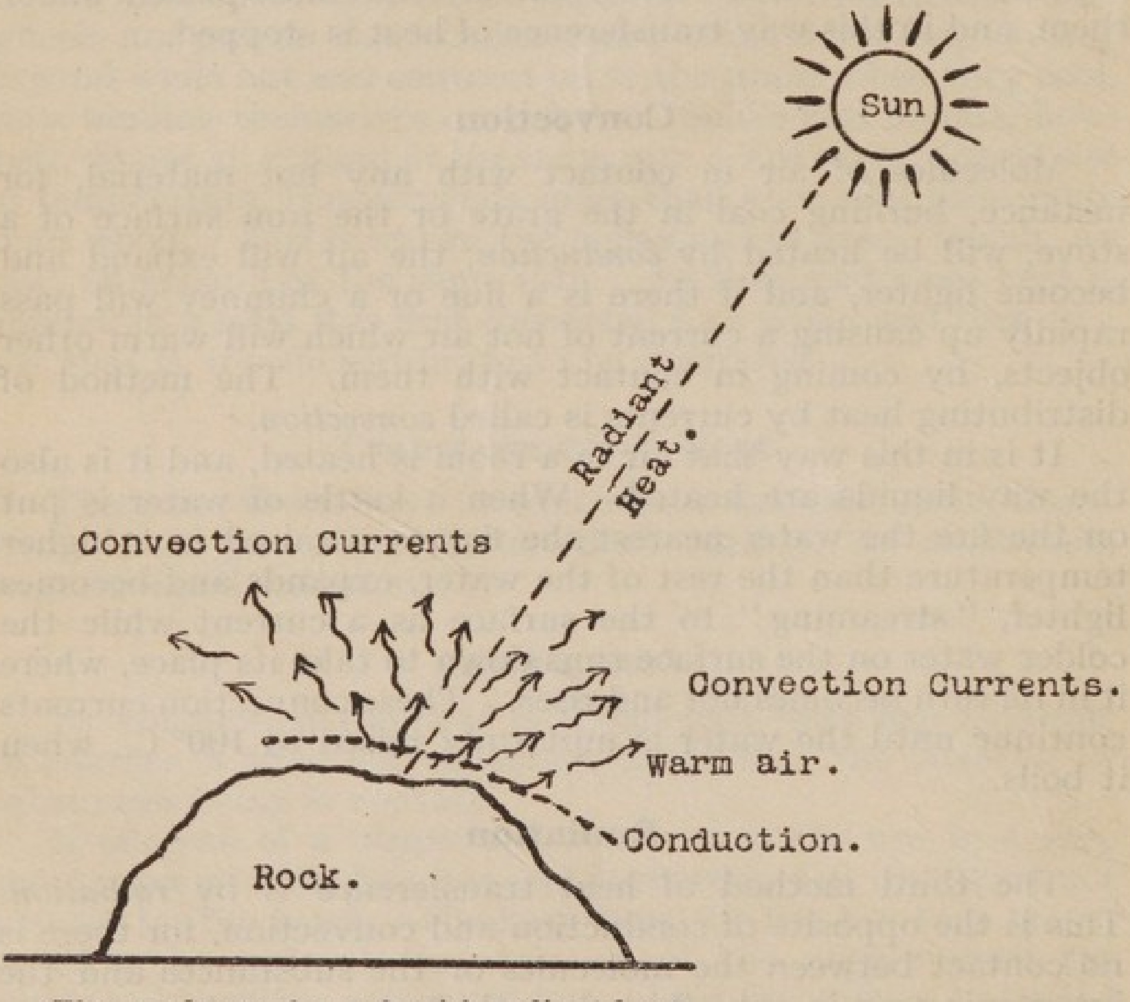
The sun is the chief source of radiant heat, and warms the objects which it strikes on the earth's surface, without warming anything in between.

It is supposed that there are heat waves just as there are sound and light waves, and it is only when these waves impinge on some receiver that the nature of the waves is appreciated. In some deaf people the drum of the ear is destroyed, therefore a receiver does not exist and the waves of sound pass unheard. Similarly, if the retina of the eye is destroyed there is no receiver for the rays of light to strike, and therefore the individual will not be able to distinguish daylight from darkness. The heat waves act in the same manner, until they strike the earth they are not appreciated as heat.

These waves of radiant heat whether given off by the sun, a fire, or a stove, unlike the waves of the ocean due to the wind, pass out in all directions.

The air is heated indirectly by its contact with the warmed earth and convection currents carry the layer of warm air away and replace it by cool air.

DIAGRAM ILLUSTRATING THE THREE METHODS OF HEAT TRANSFERENCE



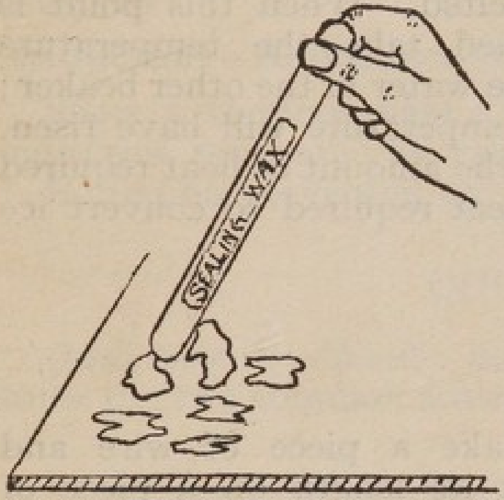
The sun heats the rock with radiant heat.
The rock heats the layer of air resting on it by conduction.
The warm air is carried off and heats other objects by convection.

Substances vary in the amount of radiant heat they absorb or give off.

The blue slates on the roof may become so hot that one can hardly touch them with the hand, while the whitewashed side of the house may feel quite cold. A dull black surface absorbs and gives off more radiant heat than a polished metal surface which reflects it. White colours are worn in the tropics because they reflect the heat, whereas black colours absorb the heat and would be uncomfortably warm to wear.

To prevent a hot substance losing its heat it should be surrounded by a non-conducting substance, placed in a vacuum to prevent convection currents and enclosed in a polished metal container whose sides will reflect any radiant heat. It is on these points that the efficiency of the thermos flask depends.

PRACTICAL CHEMISTRY—CHAPTER 7



Expt. 35

Energy

Rub a piece of vulcanite or a stick of sealing-wax with a piece of flannel.

Note that they become hot and that they will pick up small pieces of paper.

The energy expended in rubbing has been converted into heat and into electricity.

Heat

Expt. 36

Degrees

Show various forms of thermometers.

Expt. 37

Quantity

Take a $\frac{1}{4}$ lb. weight or metal body of approximately the same weight and weigh out an equal weight of water.

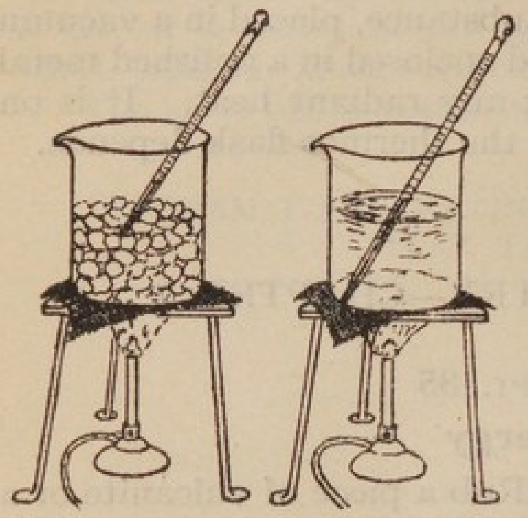
Pour the water into a cup and drop into it the metal body. Heat over a flame until the water commences to boil. The water and the metal will be at the same temperature, namely, 100° C.

Empty the boiling water into a cup half-full of water and the metal into another cup with the same amount of water. Take the temperature of the two samples of water and note that the one to which boiling water has been added is hotter than the one to which the metal was added.

Water therefore contains a greater quantity of heat than the metal.

Expt. 38

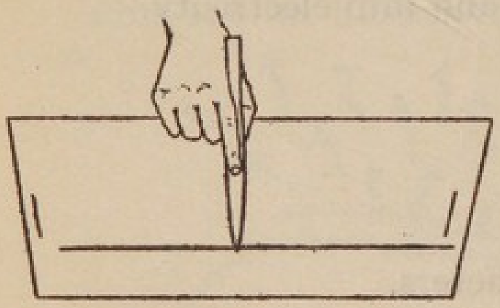
Latent Heat



Take two beakers or cigarette tins, fill one with broken-up ice and the other with water that has been cooled to 0° C. by being surrounded with ice. Heat both at the same time with a Bunsen or spirit lamp and note that the temperature of the ice will remain at zero until all the ice is melted. When this point is reached take the temperature of the water in the other beaker; its temperature will have risen, and the amount of heat required to do this is equal to the latent heat required to convert ice to water.

Expt. 39

Expansion of Metals



Take a piece of wire and mark its length on a piece of paper. Heat the wire and hold it by means of forceps over the marks and note that it is now longer.

As it cools it will contract to its original length.

Expt. 40

Conduction

Heat one end of an iron rod in a flame and note how quickly the end held by the hand becomes hot.

Similarly strike a match and note it can be held until almost completely burnt.

Iron is a good conductor of heat and wood a bad conductor.

Expt. 41

Convection

Drop a crystal of permanganate in a beaker of water and heat; note convection currents, these are currents of water at different temperatures, the hot rising to the surface and the cooler sinking to the bottom.

Expt. 42

Radiation

Take an empty biscuit tin or kerosine oil tin, polish one of the internal faces, blacken another with blacking, and whitewash a third.

Place a spirit lamp in the centre of the tin and note the difference in temperature of the different sides by placing the hands on the outside. Black will be hottest, white next and the polished face least hot.

Expt. 43

Conduction—Radiation—Convection

Show a thermos flask; loss of heat by conduction and convection prevented by vacuum.

Radiation prevented by silvering the glass.

QUESTIONS

Question 1.—What does a thermometer measure? Name the two common scales and state how they were arrived at.

Question 2.—Describe a thermometer and name the two liquids most commonly used in their manufacture, giving the reason why they are in every respect "suitable."

Question 3.—What is "Latent heat"? Describe two experiments proving its existence.

Question 4.—What are the different methods by which heat can be transferred from one object to another? Illustrate your answer by quoting experiments.

CHAPTER 8

THE EFFECTS OF PHYSICAL FORCES

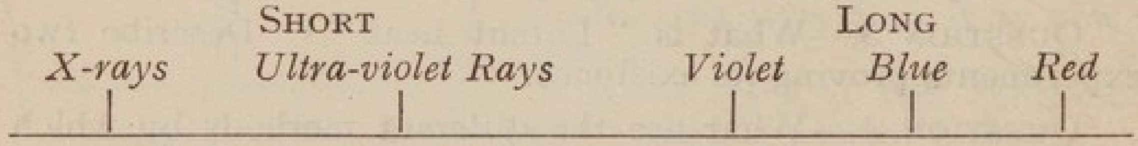
Light

LIGHT is a form of energy which is closely related to, and usually accompanies, radiant heat.

The sun gives off waves of radiant heat and running parallel to these are waves of light, which are shorter and closer than the heat waves.

Waves are measured from crest to crest, and according to the distance of " wave-length " depends the nature of the energy, whether heat, light, or electricity. For example, the wave-length of wireless is longer than that of heat and light. But light, as has been shown in Chapter 2, can be broken up into its spectrum and the colours of the spectrum each have their own wave-length, the red light has a longer wave-length than the blue, and the blue light longer than the violet. Beyond the violet there are waves of invisible light called the ultra-violet rays, and beyond these with the shortest wave-length of all come the X-rays.

Wave-Lengths



Light, unlike radiant heat, is visible and bodies that give off light are said to be *luminous ;* for example, the sun, lamps of all sorts, and substances raised to white heat.

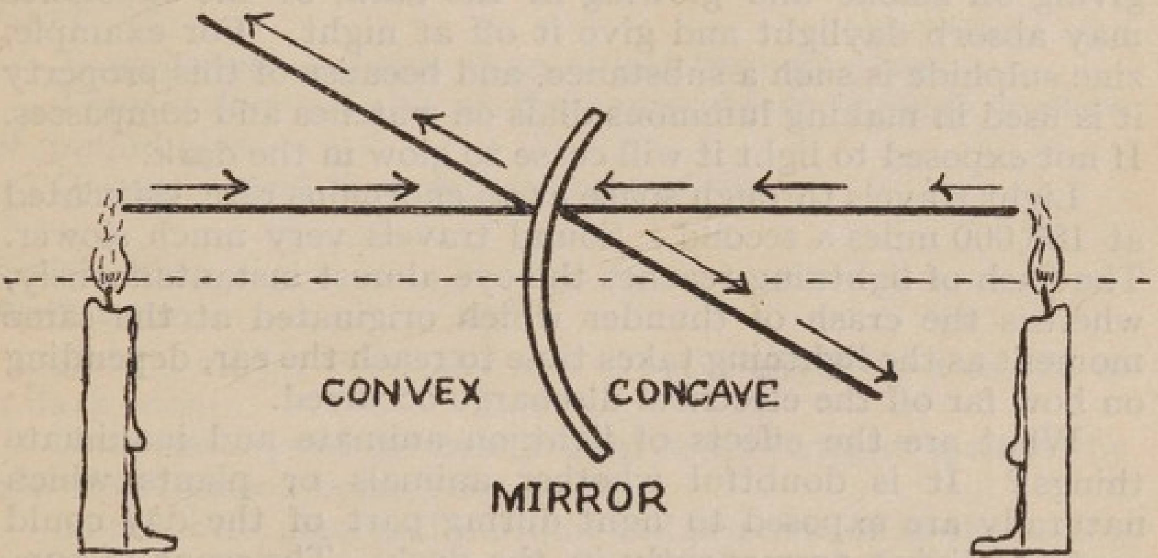
Transparency, Translucency, and Opacity

Substances may be classified according to whether or not they allow light to pass through them. Those that allow complete passage of light, for example glass, crystal, and water, are called *transparent*, and objects viewed through them are clearly visible to the naked eye. Other substances, such as ground glass, oil, etc., allow only a fraction to pass and objects appear as indistinct shadows. These substances are called *translucent*. When eventually all light is kept from passing through the substance it is said to be *opaque*. Stone, wood, and coal all come under this category.

A substance that is transparent to ordinary light may be opaque to ultra-violet light. It is said that ordinary glass is such a substance, and that ultra-violet rays will pass through quartz glass.

Reflection of Light

When a ray of light strikes a polished surface it is *reflected*, and the direction in which it is reflected depends on the shape of the surface. If it is a *plane* or flat surface the ray of light will be reflected at the same angle as that at which it strikes the surface. A ray which strikes the surface at right-angles will be reflected straight back to its source along the line it entered.



Use is made of these facts in signalling with the helio; the sun's rays are caught on a plane mirror, which is tilted to the required angle, so that the rays can be directed to an observer many miles away.

If the surface is *concave*, that is, hollow like a saucer, parallel rays are reflected inwards, whereas parallel rays are reflected outwards when they strike a *convex* or round surface.

Refraction of Light

When rays of light pass through a transparent substance, whether it is solid like glass, liquid like water, or gaseous like air, they are said to pass through a *medium*. So long as they remain in the same medium they will continue in a straight line, but should they pass from this into a denser medium, for example, from air into water, the rays are bent or *refracted*. The same thing occurs with rays passing from a dense to a less dense medium. This refraction of light is seen when a stick is partly immersed in water. The rays of light from the part above the water come to the eye unaltered, because the medium is air only, but the rays of light coming from the submerged part run straight so long as they remain in the water, but when they reach the surface they enter the less dense medium, air, and are bent or refracted, giving the stick the appearance of being crooked.

Properties and Effects of Light

Some substances have the power of storing light in their mass. This power may be the result of a chemical action such as occurs if phosphorus is exposed to the air, the phosphorus giving off smoke and glowing in the dark, or the substance may absorb daylight and give it off at night. For example, zinc sulphide is such a substance, and because of this property it is used in making luminous dials on watches and compasses. If not exposed to light it will cease to glow in the dark.

Light travels through space at an enormous rate, estimated at 186,000 miles a second. Sound travels very much slower. The flash of lightning reaches the eye almost instantaneously, whereas the crash of thunder which originated at the same moment as the lightning takes time to reach the ear, depending on how far off the electrical discharge occurred.

What are the effects of light on animate and inanimate things ? It is doubtful whether animals or plants which naturally are exposed to light during part of the day could continue living permanently in the dark. The want of sunlight is recognized as one of the causes of stunted and unhealthy children, and flowers deprived of it grow lanky and colourless, and are unable to absorb carbon dioxide, on which they live. Sunlight tans the skin and bleaches the hair, and actually in some animals, such as the chameleon, reacts on the pigment of their skin changing the colour almost instantaneously from one tint to another, as the animals move from light to shade.

Apart from these health-giving properties, sunlight kills microbes and therefore helps to purify the air, the ground, and to some extent the water.

On certain chemical substances it has a very definite action, so much so that they have to be kept either permanently in the dark or in dark tinted bottles. The action of light on some of the salts of silver is made use of in photography. The photographic film or plate is coated with silver bromide, and when an exposure is made the rays of light of varying intensity convert the silver salt into black metallic silver—the degree of blackness depending on the brightness of the rays passing in. When the film or plate is developed, these features are brought out and the film obtained is called the " negative," because the lights and shades are the opposite of what actually existed in the scene photographed.

Hydrogen and chlorine, two gases, can be mixed together in the dark without danger, but if the mixture is exposed to sunlight a violent action takes place resulting in an explosion.

PRACTICAL CHEMISTRY—CHAPTER 8

Light

EXPT. 44

Spectrum

Repeat Expt. 7.

EXPT. 45

Transparency.—Look at an object through glass.

Translucency.—Look at the same object through a piece of ground glass.

Opaque.—And then with a sheet of tin.

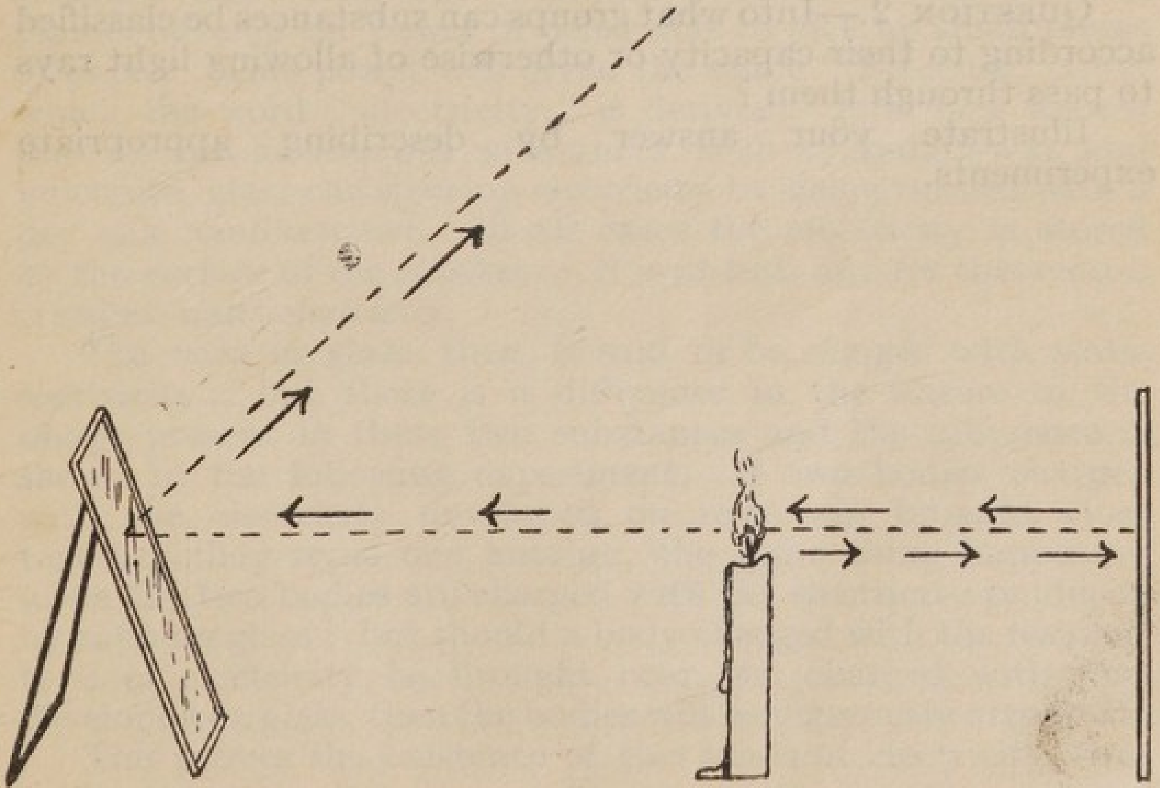
EXPT. 46

Reflection

Hold a mirror vertically and at right-angles to a candle in a dark room.

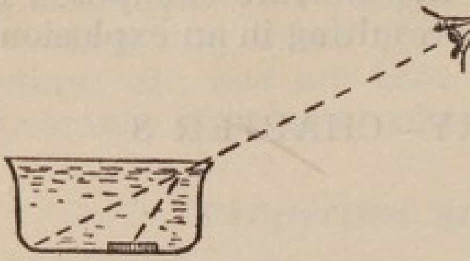
Note the rays of the candle light come straight back to the eye of the person behind the candle.

Now tilt the glass and the light will be reflected to the ceiling or floor, depending on the angle and direction of the inclination.



Expt. 47

Refraction



Place a penny in a saucer so that it is just not visible, now pour in some water gently and the penny will come into view, due to the rays of light being refracted by the water.

Hold a pencil upright in the water and note that it appears bent, due to refraction.

Expt. 48

Effects of Light

Dip a piece of filter paper in some silver nitrate solution.

Expose the paper to light and then wash the paper in a salt solution.

Notice the slow blackening of the paper due to exposure to light. (This may take some hours.) This experiment can be repeated in the dark to show that the blackening then does not occur.

QUESTIONS

Question 1.—Define the term "Light." What happens to a ray of light when it strikes and penetrates a thick transparent glass plate ?

Question 2.—Into what groups can substances be classified according to their capacity or otherwise of allowing light rays to pass through them ?

Illustrate your answer by describing appropriate experiments.

CHAPTER 9

THE EFFECTS OF PHYSICAL FORCES—ELECTRICITY

No attempt is made in this chapter to describe anything but the most elementary facts about Electricity. The subject is so large that to treat it thoroughly would require not only many chapters but many volumes, and then it would be a tale only half told.

The first question, " What is Electricity ? " must remain unanswered for no one can define it. It is undoubtedly a force or energy and resembles heat in many respects, for like the latter it is invisible yet can be felt, it can be stored in substances, it can pass along certain metals which are generally good conductors of heat, and finally, the wireless waves resemble those of radiant heat, in that they pass through the air unnoticed and only become apparent when they impinge on some receiver, the aerial wire in the case of wireless and in the case of radiant heat some object.

Static Electricity

The first realization that some peculiar force existed in certain substances was three hundred years ago, when a scientist discovered that resinous substances rubbed with flannel develop the power of being able to pick up small pieces of paper. This power was given the name " Electrics," from which the word " electricity " is derived. It is now known that as well as resinous substances, such as sealing-wax and vulcanite, glass can develop electricity by being rubbed with a dry silk handkerchief. In all cases the electricity is stored on the surface of the substance, it is at rest, and for this reason is called *static electricity*.

The resin or glass, then, is said to be *charged* with static electricity ; but there is a difference in the nature of the *charge* present in these two substances and the difference is shown in the following experiment. If two bodies charged with the electricity developed on resin are brought close together they repel one another, the same thing happening when the two bodies are charged with the electricity produced by rubbing glass ; but should a body charged with the resinous type of electricity be brought near one charged with that developed on glass, then the bodies will be vigorously attracted.

This proves the existence of two kinds of electricity, that in glass is called *positive,* and that in sealing-wax is called *negative.* It is a law therefore in static electricity that "*like charges repel*" one another—for example, two positive charges or two negative charges—and *unlike charges attract*—a positive charge attracts a negative charge.

When electricity is allowed to escape from a body that body is said to *discharge* its electricity, and if the discharge is caused by touching the body with the hand then it is felt as an electric shock, the severity of which will depend on the charge of electricity, and may be sufficient to kill the individual.

Conductors and Insulators

It has been shown that metals are amongst the best conductors of heat, and they are likewise found to be good *conductors of electricity,* for they allow an electric charge to flow through them and so escape. If a metallic body is charged and placed on the ground, all the electricity will be conducted to the earth and be lost, but if such a substance as dry glass or vulcanite is interposed between the body and the ground, then the charge is retained.

Substances which prevent the loss of electricity are known as *insulators.*

Magnetism

Before dealing with other properties of electricity it would not be out of place here to describe *magnetism,* because this force has many things in common with static electricity.

There exists in nature a mineral ore called magnetite or lodestone, containing an oxide of iron which possesses the power of attracting iron filings. This property is called *magnetism,* and can be transferred to steel by stroking it with this mineral ore. These magnetized pieces of iron are known as *magnets* and may be in the form of straight bars or shaped like a horseshoe, for in this latter form they are most powerful. If a bar magnet is suspended, so that it lies horizontally and can swing freely it comes to rest always in the same position, one end pointing north and called the North Pole of the magnet, the other end pointing south and called the South Pole.

A bar of iron which has not been magnetized can be converted into a magnet if it is placed in the centre of numerous coils or loops of wire through which a current of electricity is passing. This type of magnet is called an *electro-magnet,* and it maintains its magnetism only so long as the current is passing, during which time it acts in every way similarly to an ordinary magnet, attracting iron filings and swinging to take up a north and south position. The area around a magnet in which attraction of iron filings takes place is known as the *magnetic field.*

CHAPTER 9 (*continued*)

ELECTRICITY (*continued*)

Current Electricity

When a substance is charged with electricity and the charge is conducted away as soon as it is formed, an *electric current* results, and this type of electricity is known as *current electricity*, as distinct from static electricity. The difference of the two types, static and current, can be compared with water stored in a cistern and water circulating in a pipe, and as water can only run from a higher to a lower level, so a current of electricity will only run from a higher to a lower pressure or *potential*, and along the wires usually runs from the positive to the negative pole. The greater the difference of *potential* the greater the power of the current, and this power or *electromotive force* is measured in *volts*. A current of high *voltage* means a very powerful current, it denotes degree or intensity just as temperature denotes the degree of heat. The *quantity* of *electricity* is measured by the *ampere*, and this again is analogous to the quantity of heat which was measured by the calorie.

The flow of electricity passing along a wire can be partly or wholly obstructed and loses power or force in proportion to the degree of obstruction ; such obstruction is called *resistance* ; and in electrical work it is frequently required either to measure the resistance of a wire, or to introduce a wire with a known resistance.

This necessitates having some means of stating the degree of resistance arithmetically, and so a unit of resistance is used which is called the *ohm*.

It is found that substances increase in their resistance as they become worse and worse conductors. Copper wire has little resistance, especially if thick and short. Iron wire has greater resistance, water still greater, and gas and the insulating substances the greatest of all, in fact, their resistance is complete and no current can pass.

How is a current of electricity produced ? The simplest method is by means of a *voltaic cell or battery*. The most elementary form of battery consists of a glass jar filled with sulphuric acid in which are placed a zinc and a copper rod. If these are connected together with a copper wire an electric current will flow along the wire from the positive copper to the negative zinc pole, and within the battery from the negative to the positive pole, and in this way a complete *circuit* is formed. The zinc is gradually dissolved away with the formation of zinc sulphate and the freeing of hydrogen gas. The energy resulting from this chemical action takes the form of electricity.

The bubbles of hydrogen collect on the copper rod, and if not removed by some means will eventually weaken or destroy the current, because, as already shown, all gases resist strongly the passage of electricity. Blocking of the current in this manner is known as *polarization*, and is generally obviated by adding to the acid, or by surrounding the copper with some substance which will give up oxygen and by doing so will unite with the hydrogen to form water. There are several such substances, among which are manganese dioxide and potassium bichromate. The ends of the copper and zinc rods to which the wire is attached are known as the *poles* or *electrodes* ; the copper pole or electrode from which the current flows is called the *positive pole* or *anode*, whilst the zinc pole is the *negative pole* or *kathode*.

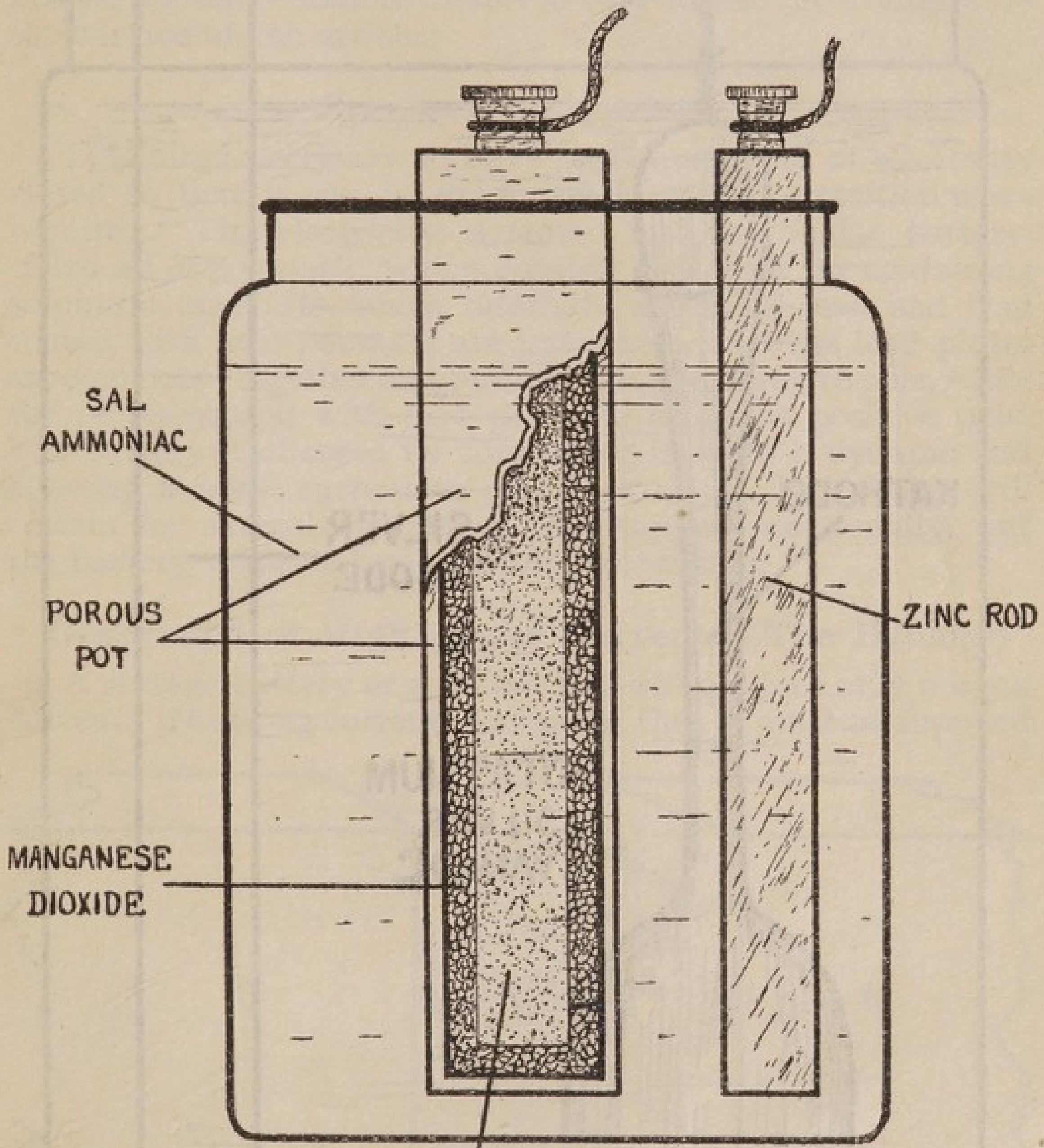
Leclanché Cell

The voltaic cell most commonly used for electric bells and telephones is the *Leclanché cell*. It consists of an outer glass jar, holding a solution of ammonium chloride (sal-ammoniac) in which stands a zinc rod, the negative pole of the battery, and a porous pot containing carbon which acts as the positive pole and which is prevented from polarizing by being surrounded by a mass of manganese dioxide. The action which takes place is as follows: The zinc dissolves in the sal-ammoniac and hydrogen gas is set free and is destroyed by the manganese dioxide. This type of cell will only give continuous currents for short periods, but it requires very little attention and will last for a long time if only used intermittently. There are many other types of cells, both dry and wet, a description of which can be found in any textbook on the subject.

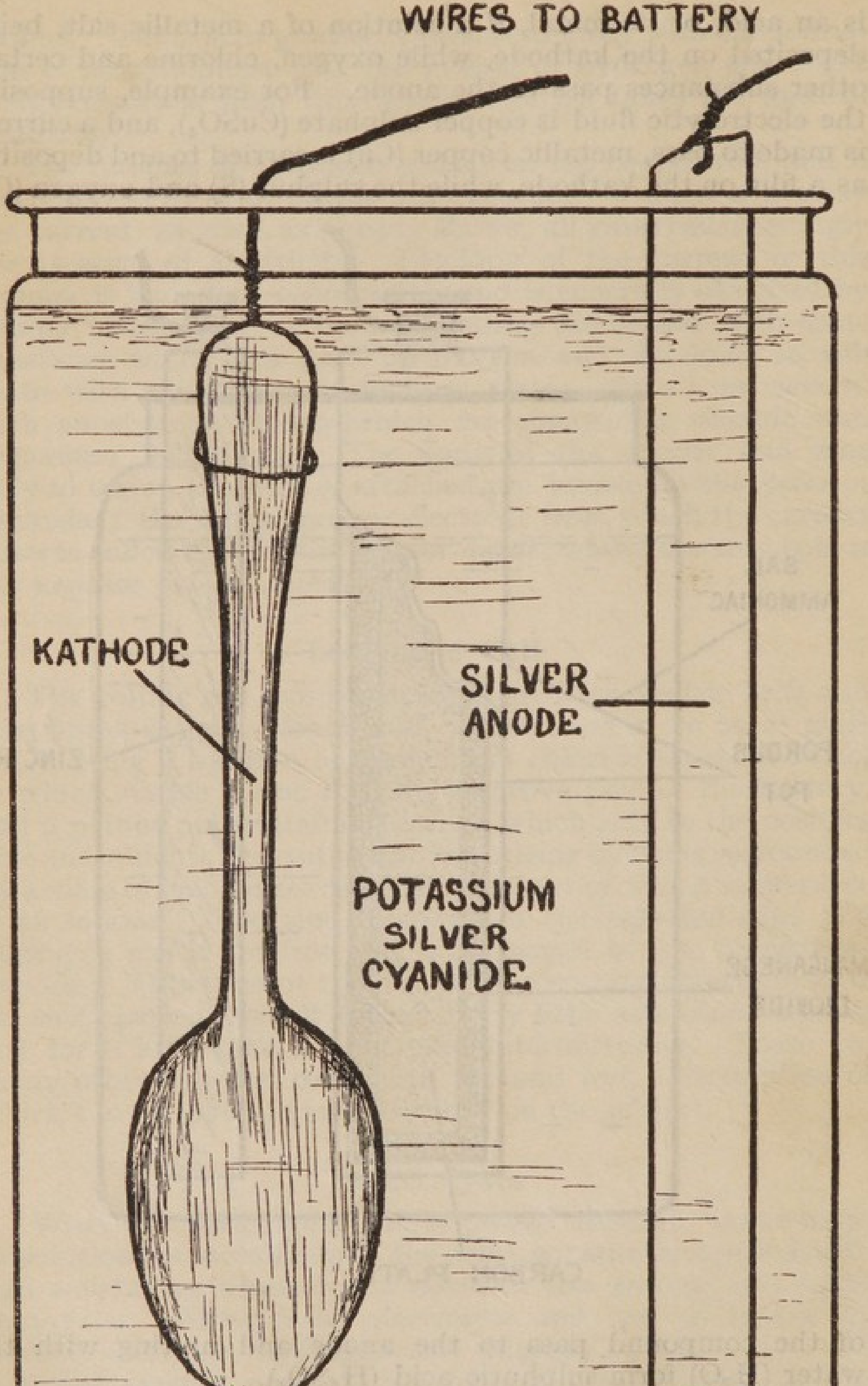
Electrolysis

When an electric current is passed through dilute acids or solutions of metallic salts the acids or salts are decomposed, and substances which break down in this manner are called *electrolytes*. Water is an electrolyte and can be broken up into its elements, hydrogen and oxygen, by the passage of a strong electric current, the process being known in this particular case as the *electrolysis of water*.

It has been shown that within the cell the current passes from the *negative* pole or *kathode* to the *positive* pole or *anode*, and the electrolytic fluid is broken up, its hydrogen, if the fluid is an acid, or its metal, if a solution of a metallic salt, being deposited on the kathode, while oxygen, chlorine and certain other substances pass to the anode. For example, supposing the electrolytic fluid is copper sulphate ($CuSO_4$), and a current is made to pass, metallic copper (Cu) is carried to and deposited as a film on the kathode, while the sulphur (S) and oxygen (O_2) of the compound pass to the anode and uniting with the water (H_2O) form sulphuric acid (H_2SO_4).



The plating or electroplating by metals occurring in this manner can be assisted if the anode consists of the metal to be plated, the electrolytic solution being a salt of that metal; for the metal will dissolve and will in this way maintain the



SILVER PLATING

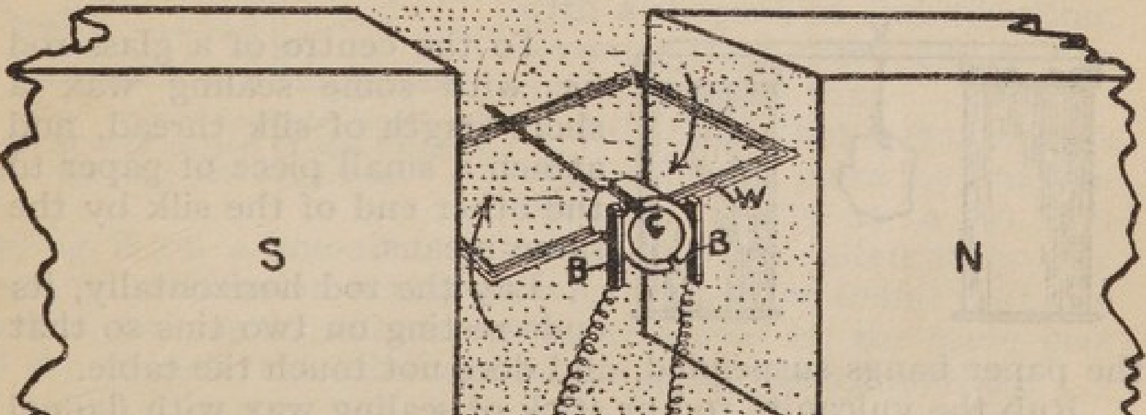
strength of the electrolyte. For this reason the anode in silver plating is silver and the solution is potassium silver cyanide ; similarly, the anode in copper plating is copper and the electrolyte, copper sulphate. The article to be electroplated is made the kathode, while the anode is formed by the plating metal, which is dissolved and deposited evenly over the surface of the article.

Storage of Electricity

It is often necessary to have large quantities of electricity stored so that it can be drawn off in small quantities when required. The electricity is stored in a particular battery, called an *accumulator*, which consists of a glass jar containing sulphuric acid into which alternate sheets of lead and lead coated with lead peroxide are immersed. All the lead plates are connected together and terminate in the negative pole, while the plates coated with lead peroxide form the positive pole. The battery is charged by connecting it up to a dynamo and allowing a weak current to pass in slowly. The charge will remain for several months, if only tapped occasionally, but the battery will eventually require recharging.

Generation of High-power Currents—The Dynamo

A storage battery or a voltaic cell will produce only a weak current ; if a strong current is required then it must be supplied from a machine called a *dynamo*, and this is now the chief source of electricity. It was shown that a bar of iron became



S and N are the poles of a magnet.
W is a coil of wire rotating through the magnet field in the direction of the arrows.
C is an arrangement (commutator) by means of which the current generated in W can be picked up by the brushes BB.
P Wire along which current flows.

magnetized if it lay inside a circuit of electricity. Now the reverse is found to occur if coils of wire are rotated through a magnetic field, then the wire picks up an electric current which will flow in which ever direction the coil of wire rotates. From a dynamo currents of very high voltage or power can be obtained.

A current that flows in one direction is called a *direct* current, but when it goes back and forward along a wire it is called an *alternating current.*

PRACTICAL CHEMISTRY—CHAPTER 9

Electricity

EXPT. 49

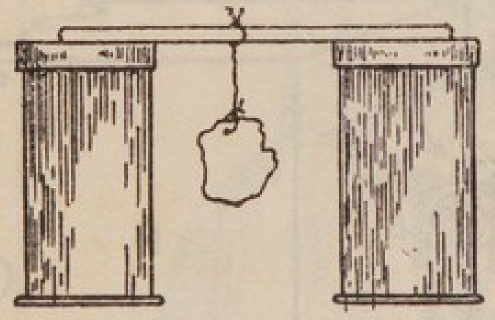
Static Electricity

Rub a stick of sealing-wax or rod of vulcanite with flannel and then a glass rod with silk.

Note that these substances have acquired the property of picking up small pieces of paper.

EXPT. 50

Positive and Negative Electricity



To the centre of a glass rod fix with some sealing wax a short length of silk thread, and attach a small piece of paper to the other end of the silk by the same means.

Lay the rod horizontally, its ends resting on two tins so that the paper hangs suspended, and does not touch the table.

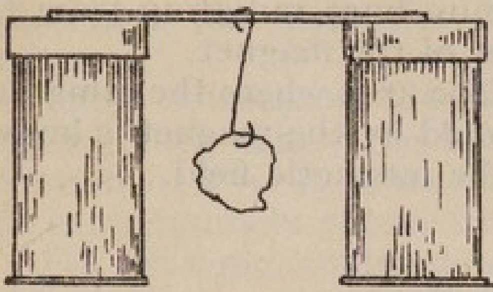
Rub the vulcanite rod or stick of sealing wax with flannel and keep it in contact with the paper for a short time in order to charge the latter with negative electricity.

Pull the vulcanite rod or sealing-wax away from the paper and then bring them near each other again; the paper is repelled because it has the same charge as the wax, in other words, *like charges repel.*

Now rub a glass rod with silk to produce positive electricity, bring it near the paper which has been negatively charged with the wax, the paper will be attracted because *unlike charges attract.*

EXPT. 51

Conductors and Insulators of Electricity



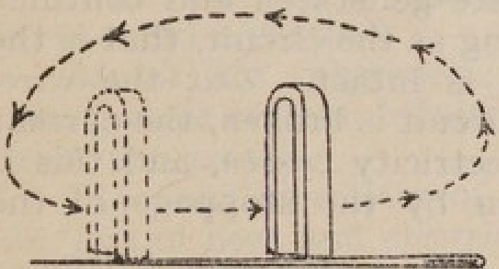
The previous experiment should be repeated, substituting thick and thin copper wire (a strand of flex) for the glass rod and silk for the suspension of the paper.

The paper will always be attracted because it never retains its charge, the metal wires allowing the electricity to leak away.

These two experiments prove that glass is a non-conductor or insulator and metal a conductor of electricity.

EXPT. 52

Magnetism



The arrows indicate needle is stroked with magnet in one direction only, the same pole being used.

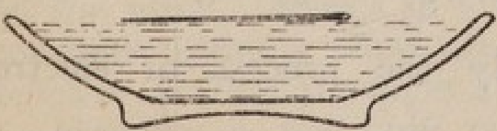
(i) Show that iron is attracted by a magnet.

(ii) Show that a piece of steel (darning needle) when stroked with a magnet in one direction, becomes magnetized, and will in its turn attract iron.

(iii) Show that a magnet has two different poles by magnetizing a needle as in (ii), and laying it on a smooth surface (glass), and alternately presenting each end of a magnet when it will be found that one end of the needle will always be attracted by the same pole of the magnet, the other end of the needle being repelled.

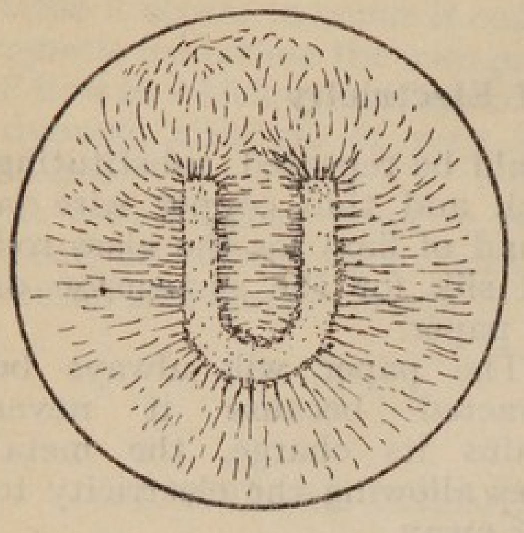
EXPT. 53

North and South Pole



Float a small and greasy magnetized needle on the surface of water in a *china* saucer. It will always take up a position roughly north and south, unless pieces of iron are near it.

Expt. 54



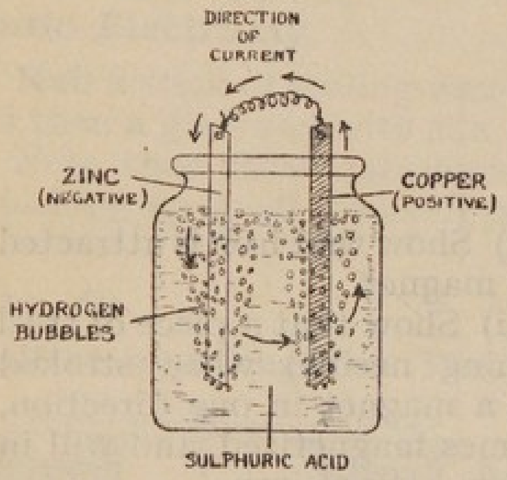
Magnetic Field

Sprinkle some fine iron filings on a sheet of white paper laid on a magnet. Note that the filings arrange themselves into definite lines radiating from the poles of the magnet.

The area where the filings are affected by the magnet is known as the magnetic field.

Expt. 55

A simple Voltaic Cell



Place a sheet of zinc and a sheet of copper in a jam pot containing dilute sulphuric acid, and connect the metals by a piece of copper wire. Bubbles of gas are generated and continue so long as the circuit, that is the wire, is intact. Cut the wire, the circuit is broken, the current of electricity ceases, and this is shown by the stoppage of the gas.

If a Leclanché cell, an accumulator or a dynamo, can be obtained, these should be shown to the student.

QUESTIONS

Question 1.—What do you understand by positive and negative electricity, and how can you prove their existence ?

Question 2.—Copper is a good conductor of electricity. Explain this and give examples of bad conductors.

Question 3.—Define the difference between static and current electricity. Name the various methods of producing these two types of electricity.

Question 4.—Describe a simple voltaic cell. How can you tell that the current of electricity is flowing ? What is polarization of the cell ?

Question 5.—Describe briefly the process of electroplating.

Question 6.—How can a piece of steel be converted into a magnet and when magnetized what properties will it possess ?

CHAPTER 10

THE ELEMENTS

As already mentioned in Chapter 1, there are a number of substances the molecules of which are composed of similar atoms.

These substances are called *elements*, and, as the name implies, they cannot be broken up further. It is as a combination of two or more of these elementary bodies that all the compounds in nature exist.

For example, common salt (NaCl) is a compound and can be broken down by chemical action into sodium and chlorine, but the reduction cannot proceed beyond this. The atoms in the molecules of sodium are all pure sodium, and similarly the atoms in the molecules of chlorine are all pure chlorine. Sodium and chlorine are therefore elementary substances.

Metals and Non-metals

Elements can be divided roughly into two main groups: *metals* and *non-metals*.

Metals are substances, such as iron, copper, gold, etc., which have certain definite properties which distinguish them from non-metals.

The chief character of metals is that they are good conductors of heat and electricity, while non-metals allow neither heat nor electricity to pass.

Another important point of difference is the reaction of the tarnish which occurs on exposing metals and non-metals to damp air. The metallic tarnish is always alkaline, whereas that of non-metals is acid. The third distinction between metals and non-metals depends on the fact that the former can be liberated at the kathode of an electric cell.

In appearance metals are opaque, their surface, if smooth, shines with a " metallic lustre," and the majority of them are insoluble in water but soluble in certain acids.

All metals, if heated sufficiently, melt, the temperature required to do this varying with the kind of metal. If a metal is pure its temperature at its melting point will always be the same. Metals in a molten condition can be made to combine, forming what are known as " alloys," for instance, brass is an alloy of copper and zinc.

Mercury is the only metal which is liquid at ordinary temperatures.

In the case of some of the elements, their presence can be detected when their salts are burnt in a bunsen flame, which is coloured yellow with salts of sodium, lilac with salts of potassium, and green with salts of barium, etc.

QUESTION

QUESTION—Define the term "element." Into what categories can all elements be divided and describe the characteristics of each ?

CHAPTER 11

CHEMICAL SYMBOLS—EQUATIONS AND VALENCY

To obviate the necessity of having to write the names of the elements and their compounds in full, certain abbreviations or *symbols* have been selected. These symbols may consist of the first letter of the Latin name of the element, or in some cases the first letter and one other.

Some examples of these are :—

Boron = B
Carbon = C
Oxygen = O
Sulphur = S
Hydrogen = H
Silver = Ag
Lead = Pb
Gold = Au
Chlorine = Cl

These chemical symbols represent one *atom* of the substance.

To represent a *molecule*, a number which shows how many atoms the molecule contains must be added to the symbol. This is done by placing the necessary number after the symbol and at a lower level.

Thus one molecule of *hydrogen* contains two atoms of hydrogen and would be represented as H_2 ; similarly, a molecule of *oxygen* contains two atoms of oxygen and would be written O_2 ; while *phosphorus*, having four atoms in its molecule, would be represented as P_4, and so on.

In the case of compounds, where there are two or more elements present, the symbol for each of the elements is placed side by side, and the number of atoms that compose their molecules is added as described above.

For example, *common salt, sodium chloride*, is a compound of one atom of *sodium* and one atom of *chlorine*, and is represented by NaCl. Water is a compound of two atoms of *hydrogen* and one atom of *oxygen*, and is shown as H_2O.

The symbols of a compound are termed its *formula*. NaCl is the formula of sodium chloride.

When it is necessary to denote the presence of more than one molecule of a substance, the particular number is placed directly in front of the formula, large figures being used. Two molecules of sodium chloride are shown as 2NaCl.

By means of chemical symbols and formulæ the action of one chemical substance on another can be shown. The action is put in the form of an *equation*, and like all equations, the two sides forming it must balance or be equal to one another. For example, if one wished to construct a simple arithmetical equation from $8 + 4$, one would select the single number 12, or two or more numbers which when added together give 12, and place them on the opposite side, separating the two sides by the sign = which represents " equal to."

$$\begin{aligned}\text{Thus } 8 + 4 &= 12\\ 8 + 4 &= 6 + 6\\ 8 + 4 &= 4 + 3 + 5\end{aligned}$$

Similarly, in chemical equations, the substances resulting from a chemical action must balance or be equal to the substances causing the action. For example, *hydrogen gas* and *chlorine gas* unite to form *hydrochloric acid*, and the reaction is shown in symbols by equation :

$$H_2 + Cl_2 = 2HCl$$
$$(2H + 2Cl) = (2H + 2Cl)$$

Similarly, *caustic soda* (NaOH) will be acted on by *carbonic acid gas*, forming *sodium carbonate* (Na_2CO_3) and water (H_2O). Put in the form of an equation this appears as :

$$2NaOH + CO_2 = Na_2CO_3 + H_2O$$
$$(2Na + 4O + 2H + C) = (2Na + 4O + 2H + C)$$

But if written in full would read :

Two molecules of caustic soda plus one molecule of carbon dioxide gas produce one molecule of sodium carbonate and one molecule of water.

The symbolic equation conveys to the mind of the trained chemist as much as the same equation written in full, and it does so more quickly and more concisely.

It is difficult at first for the beginner to understand why the atoms of the elements composing a chemical substance are always present in a definite number. Why, for instance, should water have a constant formula H_2O, that is two atoms of hydrogen and one of oxygen, and not the other way about ?

This introduces the subject of *valency*, which may be described as the *combining power of the atoms of one element with the atoms of other elements.* Some elements have greater combining power than others, for example, such elements as *sodium*, *chlorine*, and *silver* can only combine with one atom of any other element. Their power of combining is said to be 1, and they are termed *monovalent* elements. Chlorine and sodium are both monovalent elements, only a single atom of each combining; therefore, the formula of the compound sodium chloride, which results from this union, would be NaCl.

Now if hydrogen is taken to represent one of the monovalent elements, it will be found that elements like *oxygen, calcium, copper*, and *zinc*, etc. would not form compounds unless 2 atoms of hydrogen were present. Their combining power is double that of hydrogen and therefore they are classed as *divalent* elements.

It requires 2 atoms of the monovalent element hydrogen to combine with 1 atom of the divalent oxygen to form the compound water.

$$H_2 + O = H_2O$$

Other elements need 3, 4, 5, and more atoms of a monovalent element before a compound can be formed; they would be called, *trivalent, tetravalent*, and *pentavalent* elements. Nitrogen requires 3 atoms of hydrogen to form *ammonia* (NH_3) and is therefore a trivalent element. *Carbon* is tetravalent, requiring 4 atoms of hydrogen to form the compound *methane* (CH_4).

If a divalent element combines with another divalent element, only 1 atom of each will be present in the compound formed, and similarly, if the elements are both trivalent, only 1 atom of each enters into their combination.

Calcium and *oxygen* are divalent substances, 1 atom of each combining to form the compound *calcium oxide, quicklime.*

$$Ca + O = CaO$$

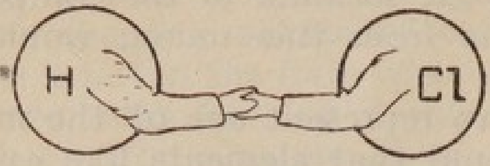
And yet each would require 2 atoms of monovalent hydrogen or chlorine to combine

$$\begin{aligned} Ca &+ 2Cl = CaCl_2 \\ O &+ 2H = H_2O \end{aligned}$$

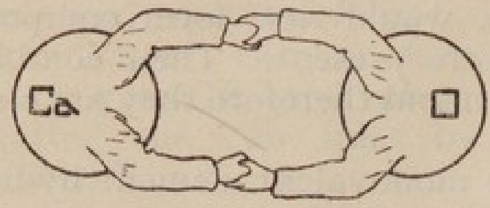
To make the meaning clearer, suppose there exist one-armed, two-armed, three-armed and four-armed men, representing in that order, monovalent, divalent, trivalent, and tetravalent elements.

Should two one-armed (monovalent) men meet and clasp hands, neither will possess a free hand (Fig. I); similarly, if two double-armed (divalent) (Fig. II) men clasp hands all their hands will be engaged, and this will also occur if two three-armed (trivalent) men or two four-armed (tetravalent) men meet.

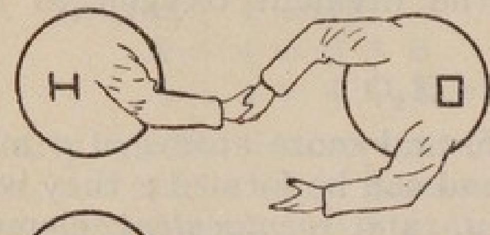
But should a one-armed meet a two-armed man only one hand of the latter will be clasped, the other would be free (Fig. III), and the two men could not therefore combine in the full action of hand-clasping, which could only be effected by



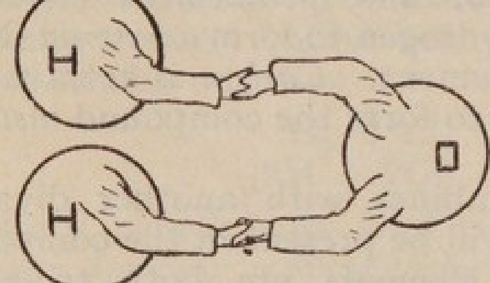
= HCl FIG. I



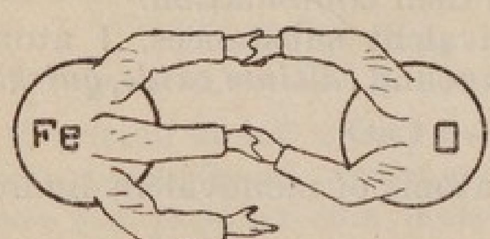
= CaO FIG. II



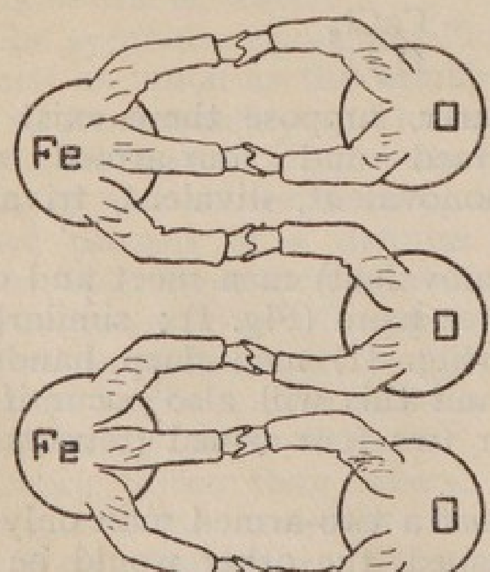
= NO UNION FIG. III



= H_2O FIG. IV



= NO UNION FIG. V



= Fe_2O_3 FIG. VI

calling another one-armed man to clasp the empty hand (Fig. IV). The divalent element oxygen therefore requires two atoms of the monovalent element hydrogen before union can take place. Similarly, with three-armed and four-armed men, three and four one-armed men would be required to make the hand clasping complete.

The trivalent element nitrogen requires three atoms and the tetravalent element carbon requires four atoms of the monovalent element hydrogen ; in symbols, NH_3 and CH_4,

Now what happens when a three-armed (trivalent) man meets a two-armed (divalent) man ? They cannot combine as they are (Fig. V), and more men of each kind must be brought along so that an equal number of hands exists on each side. It will require three two-armed men and two three-armed men, that is six arms in all, to effect a union of these two different types (Fig. VI).

Iron (Fe) is trivalent, oxygen (O) is divalent, and therefore before they will combine, there must be, as in the case of the men, two atoms of iron and three of oxygen, expressed in symbols, Fe_2O_3

The hand-clasping of a four-armed man with a three-armed man is also impossible, the same balance of hands is necessary on each side, and the least number of hands to effect this would be twelve, that is, four three-armed men on one side and three four-armed men on the other.

QUESTIONS

Question 1.—Show how a chemical equation can be represented by symbols.

Explain the reason for each letter and figure in the example you give.

Question 2.—What do you understand by the term "valency"?

Give examples of monovalent, divalent and trivalent elements.

CHAPTER 12

COMPOUNDS—ACIDS AND BASES AND THEIR PROPERTIES

Chemical Compounds

WHENEVER two or more elements combine together they form a *chemical compound*, and these compounds in turn may combine producing other substances. There are therefore existing in nature an immense number of different compounds, and the chemical action which occurs in producing these may be one of three distinct types.

The Union of Two Elements.—The simplest form of chemical reaction is where one element acts on another. It is obvious the compound resulting from this action can only contain two elements. The form it takes may be solid, liquid, or gaseous; for example, when most metallic elements are exposed to the air, or heated in the presence of air, compounds which are called *oxides* are formed with the element oxygen. Iron "rusts" when it is exposed to the air—that is, its surface is covered with a compound of iron and oxygen known as *iron oxide.*

$$2Fe + 3O = Fe_2O_3$$

Copper when heated in the presence of air takes up oxygen and forms a compound, *copper oxide.*

$$Cu + O = CuO$$

If the element carbon is burnt in oxygen, a gaseous compound is formed called *carbon dioxide gas.*

$$C + O_2 = CO_2$$

If the gases hydrogen and oxygen are mixed together and ignited by an electric spark, they combine together to form a liquid, *water.*

$$H_2 + O = H_2O$$

The Union of an Element and a Compound.—Secondly, there is the action where one element combines with a compound. If the element sodium is placed in the compound water, a new compound, *sodium hydroxide,* is formed, hydrogen being set free at the same time.

$$2Na + 2H_2O = 2NaOH + H_2$$

Most metals if mixed with dilute acids form compounds with the liberation of hydrogen gas. For example, the metallic element zinc when added to the compound sulphuric acid produces *zinc sulphate* and *hydrogen gas*.

$$\underset{\text{Zinc}}{Zn} + \underset{\text{Sulphuric acid}}{H_2SO_4} = \underset{\text{Zinc sulphate}}{ZnSO_4} + \underset{\text{Hydrogen gas}}{H_2}$$

The Union of Two Compounds.—Lastly, we have the reaction between two compounds when two or more new compounds are produced. For example, when the compound hydrochloric acid acts on the compound sodium hydroxide, there is an interchange of elements, the hydrogen of the acid uniting with the hydroxyl part (OH) of the sodium compound to form *water* and the sodium combining with the chlorine of the acid to form *sodium chloride*.

$$NaOH + HCl = NaCl + H_2O$$

If chalk, calcium carbonate, is mixed with hydrochloric acid three new compounds are formed—a solid, *calcium chloride*, a liquid, water, and a gas, *carbon dioxide*.

$$CaCO_3 + 2HCl = CaCl_2 + H_2O + CO_2$$

Acids, Bases, and Salts

Chemical compounds may be divided into three main classes—namely, *acids, bases*, and *salts*.

ACIDS

Acids are compound substances which contain hydrogen, which hydrogen can easily be replaced by a metal, either when the acid is in contact with the metal itself or with a compound of the metal called a base. For example, hydrochloric acid (HCl) is a compound of hydrogen and chlorine, and should it come in contact with a metal such as sodium, the place the hydrogen element occupied in the acid compound is taken up by the sodium, the hydrogen gas being liberated.

$$Na + HCl = NaCl + H$$

Similarly should the acid act on a metallic base such as NaOH, the hydrogen of the acid would again be replaced by the metal element of the base.

$$\overset{\text{Base}}{NaOH} + \overset{\text{Acid}}{HCl} = \overset{\text{Salt}}{NaCl} + \overset{\text{Water}}{H_2O}$$

In this example the hydrogen liberated unites with the hydroxyl part of the metallic base to form water.

Acids are usually colourless liquids or gases heavier than water and having a sour taste. The stronger acids will corrode metals, etc., forming salts of the particular metal they act on. They also react with alkalis forming salts.

They produce definite changes in the colour of certain dyes. For example, they turn litmus solution from blue to red.

Some of the acids in common use are :

Nitric acid (HNO_3).
Sulphuric acid (H_2SO_4).
Hydrochloric acid (HCl).
Carbonic acid (H_2CO_3).

ALKALIS OR BASES

Alkalis or *bases* are compounds of a metal with OH (which is termed the *hydroxyl radicle*), the resulting compound being called a metallic hydroxide.

Sodium + the hydroxyl radical = sodium hydroxide, or caustic soda

$$Na + OH = NaOH$$

Sometimes a compound substance such as ammonium (NH_4) may act like a metal and form a hydroxide.

$$NH_4 + OH = NH_4OH = \text{Ammonium hydroxide.}$$

These alkalis, or bases, react with acids to form salts, e.g. hydrochloric acid and sodium hydroxide react, forming sodium chloride—a salt and water.

$$NaOH + HCl = NaCl + H_2O$$

There are certain metallic oxides and carbonates which are alkaline in their action, turning litmus blue and forming salts with acids.

Some examples of these are :

Barium oxide or baryta (BaO) when added to water, barium hydroxide ($BaOH$) is formed.

Calcium oxide and magnesium oxide (CaO and MgO) also form their hydroxides when added to water.

Washing soda, sodium carbonate (Na_2CO_3), although not a hydroxide nevertheless gives an alkaline reaction with litmus and forms salts with acids.

SALTS

These are compounds formed by the action of an acid on a metal or on a compound radical such as ammonium (NH_4).

$$Na + HCl = NaCl + H$$

Sodium chloride, or common salt, is formed, hydrogen being set free.

In this example the hydrogen has been entirely replaced, and the resulting salt is a *neutral* or *normal salt*. Or it may be only partly replaced, when the resulting salt will be an *acid salt*. Acid sodium sulphate ($NaHSO_4$) is an example of an acid salt, a solution of which will turn litmus red.

Salts are named from the metal and acid which produce them. The salt derived from the action of sulphuric acid and zinc is called zinc sulphate. Salts formed from the commonest acids are :

Sulphates from sulphuric acid.
Sulphites from sulphurous acid.
Chlorides from hydrochloric acid.
Nitrates from nitric acid.
Nitrites from nitrous acid.
Carbonates from carbonic acid.
Phosphates from phosphoric acid.

As a general rule it may be taken that acids ending in *ic* form salts which end in *ate*, exceptions to the rule being the acids of the halogens, namely, hydrochloric, hydrobromic, hydriodic acids, the salts of which end in *ide*, for example, chlorides, bromides, and iodides. Salts derived from acids ending in *ous*, terminate in *ite*.

There are an endless number of salts found in nature. They may occur as solids or in solution, and are collected and purified for chemical, medicinal, and commercial purposes.

Salts are usually crystalline, some being very soluble in water, such as those of sodium, potassium, and ammonia, while others are insoluble or very feebly soluble such as barium sulphate and silver chloride. Many salts crystallize out from solution when the water or other solvent evaporates.

Reaction of a Solution—Indicators

If an acid and an alkali are mixed together in such amounts that the resulting mixture contains no excess of either, the solution is said to be *neutral*.

In chemical work it is frequently necessary to form neutral solutions, either by removing the acidity by the addition of the right amount of alkali or vice versa, and it is essential to have some easy and accurate method of judging when the neutral point is reached. Substances known as *indicators* are used. These are either vegetable colouring matters or some specially prepared chemicals capable of altering their colour on the addition of an alkali or an acid. The indicators commonly in use are :

	Reaction with		
	Acid	Neutral	Alkali
Litmus	Red	Plum	Blue
Methyl orange	Red	Orange	Yellow
Phenol phthalein	Nil	Nil	Red

These indicators are made into dilute solutions with either distilled water or with spirits ; phenol phthalein is known as a " one-way " indicator, as it only gives a colour with alkalis.

The following is the method of neutralizing an alkaline solution. A little litmus solution is added to the alkali and a blue colour at once develops. An acid solution is run in gradually from a glass tube until one drop of it just turns the blue to a plum colour. The solution is now neutral, any further addition of acid will make the mixture acid and the colour will be changed to red.

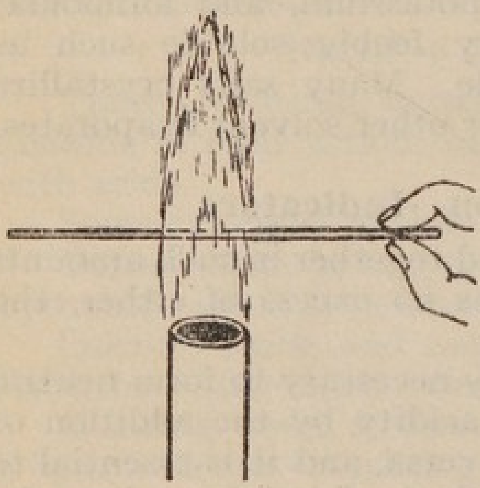
PRACTICAL CHEMISTRY—CHAPTER 12

Compounds

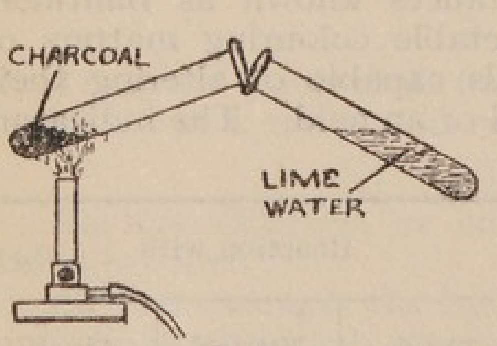
EXPT. 56

Simple

Leave a few clean iron nails in water overnight, the shiny metallic surface has disappeared, a dull red powdery substance takes its place, the iron having combined with oxygen to form " rust."



Heat a piece of bright copper in a flame ; a black substance is formed on the surface of the metal ; the copper has combined with oxygen from the atmosphere to form copper oxide.



Heat some powdered charcoal in a test-tube (see diagram) to a dull red. The carbon combines with oxygen to form CO_2.

Place the mouth of the test-tube over another test-tube or beaker containing lime water, and note the film of insoluble calcium carbonate which forms on the surface of the lime water by the union of lime and CO_2.

EXPT. 57

Compound with an Element

Repeat Expt. 28

Expt. 58

Two Compounds

Half-fill a tin lid with a strong solution of caustic soda. Add strong hydrochloric acid (HCl) (a few drops at a time) until the mixture ceases to turn red litmus paper blue. Now gently heat the solution and continue heating until all the moisture has evaporated. A white crystalline solid remains. Examine the crystals; they are cubical in shape and salty in taste, and correspond in every respect to crystals of ordinary salt.

Acids

Caution.—*When handling acids in concentrated form, care should be taken to prevent any from getting on the hands, face, or clothing.*

Always add acids to water and not water to acids.

Do not heat hydrochloric or nitric acids.

Expt. 59

Acid + Metal

To half an inch of water in a test-tube add an equal quantity of hydrochloric acid and dip in a blue litmus paper. The paper turns red immediately, indicating an acid.

Now drop two or three granules of zinc into the solution, bubbling commences and a gas (hydrogen) is given off, the zinc having displaced the hydrogen from the compound hydrochloric acid to form zinc chloride. Leave (overnight) until bubbling ceases, and drop a blue litmus paper into the solution, it turns a plum or neutral colour, showing that with the complete liberation of the hydrogen it has lost its acid properties.

Expt. 60

Acid + Base

Repeat Expt. 58. Dissolve a few of the resulting crystals in water, and test the reaction of the solution, which will be found to be neutral, the hydrogen atom of the acid has in this experiment combined with the hydroxyl radical of the base sodium hydroxide to form water; the metal sodium and the chlorine combining together to form common salt.

Expt. 61

Destructive Action of Acids

Dip a clean stirring rod into some nitric acid and touch a piece of coloured cloth, the colouring is immediately destroyed, the cloth will have become rotted and powdery when dry.

Repeat the above experiment with sulphuric and hydrochloric acids.

Expt. 62

Alkalies or Bases

Caution.—When handling do not touch with the fingers.

Dissolve a little sodium hydroxide in half a test-tube of water, dip a red litmus paper into the solution and it will turn blue, indicating an alkali.

Dissolve a few crystals of washing soda in a test-tube and test the reaction of the solution with litmus paper; it turns red litmus blue, indicating an alkali.

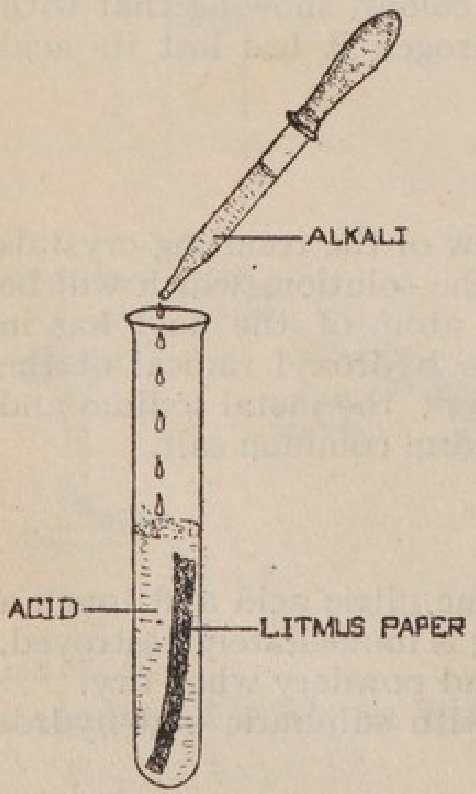
To the solution add a few drops of hydrochloric acid, and continue adding until action ceases. The gas which comes off may be identified as carbon dioxide by pouring the gas into a tube which has been rinsed out several times with lime water; the inside of the tube will assume a hazy appearance if carbon dioxide is present.

Pour the resulting solution into a tin lid and proceed as in Expt. 58. The remaining salt will be identical in character to that from this experiment and shows that sodium carbonate when treated with an acid behaves in a similar way to the hydroxide, but in addition carbon dioxide is evolved.

Expt. 63

Salts

In Expt. 60 the whole of the hydrogen was replaced, and the resulting salt when dissolved in water was found to have a neutral reaction. Dissolve a little acid sodium sulphate in water and dip a blue litmus paper in the solution; the paper turns red, showing this salt to be acid.



Expt. 64

Reaction of a Solution—Indicators

Drop a piece of blue litmus paper in a test-tube containing dilute acid, the paper turns red, now add drop by drop (with the dropper eye glass) to the tube a little dilute solution of sodium hydroxide (see diagram) and shake the tube between the drops. The paper gradually turns from red to a plum colour—this is neutral and denotes no excess of acid or alkali; continue adding the soda and the paper will turn a deep blue colour, indicating that the solution is now alkaline.

QUESTIONS

Question 1.—What is a compound ? How does it differ from a mixture ? Give examples of gaseous, liquid, and solid compounds, showing by an equation how they are produced.

Question 2.—What are the distinguishing features of acids and alkalis ?

Question 3.—How are salts formed ? Why may the different salts vary in their reaction ? Give examples of soluble and insoluble salts.

CHAPTER 13

CHEMICAL TERMS

In the chapters to follow, certain names and terms are used, the meaning of which may not as yet have been explained. It is proposed to deal with the more important of these in a separate chapter, while the others will be found defined in the Glossary.

Stable and Non-stable Substances

There are certain substances which withstand exposure to air, to heat, and to certain corroding chemicals. These are classified as *very stable substances*, because it would require extremely active and powerful agents to decompose or break them down. Other substances withstand exposure to air and to low temperatures, but decompose when treated with chemicals. These are called *moderately stable substances;* while there are others which unless preserved in a special manner decompose readily on exposure to air. These are known as *unstable substances*, and most of the terms later to be described refer to substances of this latter category.

Deliquescent or Hygroscopic Substances

There are some chemicals which have a great affinity for water, and not only do they dissolve most readily but may even absorb water from the air or from damp substances in their proximity.

They can therefore be used as drying agents and the amount of water they can absorb will depend on the nature of the substance, some continuing to absorb water until they liquefy. Such substances are called *deliquescent,* and the property they possess is called *deliquescence*. For example, common salt when exposed to the air soon becomes damp and cakes because of the moisture the crystals absorb. That is why in certain proprietary brands of table salt the manufacturers add some bone ash in order to prevent the salt deliquescing, so that it remains dry and can be sprinkled. Quicklime is another deliquescent substance, and by mixing it with bleaching powder it absorbs any moisture that would otherwise act on the bleach and decompose it.

Calcium chloride is so deliquescent that if exposed to the air for a sufficiently long period it will eventually become liquid.

Efflorescent

There are, on the other hand, substances which give up part or the whole of their water of crystallization on exposure to air, losing on doing so their crystalline form and being reduced to an amorphous powder.

These are *efflorescent* substances, of which an example is washing soda, for when its crystals are exposed to the air they become covered with a white powder, which is exactly the same substance but wanting in some of its water of crystallization.

Anhydrous

In a previous chapter it has been stated that certain substances when they crystallize out absorb water, which is known as the water of crystallization. When the crystals of these substances are heated they melt and the water of crystallization is driven off, leaving an amorphous powder which chemically is identical to the original substance, only it is without its water of crystallization.

A powder produced in such a manner is the *anhydrous form* of the substance, or the *form without water*. A number of these anhydrous substances possess the property of deliquescence, readily absorbing the water which they lost on heating.

Oxidation and Reduction

These terms are introduced here because the actions that occur are somewhat analogous to those described under Deliquescence and Efflorescent, but instead of water, oxygen is the substance which is either absorbed or driven off, and also in doing so the substance changes its chemical state.

Oxidation.—A substance is said to be an *oxidizing agent* when it readily parts with the oxygen with which it is combined. For example, hydrogen peroxide (H_2O_2) readily gives off oxygen, being transformed to water (H_2O), and therefore is an oxidizing agent.

Nitric acid and potassium permanganate are also examples of oxidizing agents of which there are quite a number.

When the oxygen given off by the agents is taken up by another substance, that substance is said to be *oxidized* or has combined with oxygen to form an oxide, and the process is termed *oxidation*.

Reduction.—On the other hand, there are chemicals which rapidly absorb oxygen, not only from the air but from its combination with elements or compounds. These substances are called *reducing agents ;* they remove the oxygen and reduce the substance to a simpler form. Carbon, or charcoal, is a reducing agent, for when heated with air it takes up oxygen to form carbon dioxide, and when heated with certain metallic oxides abstracts the oxygen, leaving the pure metallic substance. This *process of reduction* by means of carbon is employed in the abstraction of most metals from their ores.

Desiccation

A simple word for desiccation is " drying," or the removal of all moisture from a substance without altering its chemical state. This can be done by dry air sufficiently hot to assist the absorption of water yet not too hot to cause charring or ashing, or by placing the damp substance in a container (desiccator) along with a hygroscopic substance which will take up all the moisture.

Condensation, Evaporation, Distillation, Sublimation

There are a series of expressions that are used when liquids are heated or, for that matter, when some solids are heated. Some of these can be explained in the simple process of boiling water. When water is heated to boiling point it passes off in the form of *vapour* (steam) ; it is said to *vaporize*, and when the vapour finds a surface cool enough to convert it back again to water it will *condense* or *distil* over as little drops on the surface. This is called *condensation* or *distillation*. Now, it is only the pure water that vaporizes and condenses, all the impurities and salts are retained in the container which holds the boiling water. The water is said to be *distilled* water or the distillate, that is, water in its purest form containing neither gases nor salts, the latter remaining as the *residue*.

Most liquids behave in an identical manner, and distillation is the recognized method of obtaining liquid substances in their purest state. Even without the application of heat the action goes on. If water is spilt on the floor it dries up, in other words the water *vaporizes* or *evaporates*, and if there is a colder spot will condense there. The equivalent term for condensation when solids are heated is *sublimation*. The solid vaporizes and then condenses or *sublimes* as a fine powder or as crystals in a cooler part of the apparatus.

This leads one to describe the term *volatile*, which is applied to certain substances which when exposed to the air vaporize more or less rapidly.

The name is specially associated with oils and perfumes which give up their odour to the air. All the scents of flowers are due to these volatile substances.

Precipitation

In describing a chemical action between two liquid substances, when an insoluble substance results, one speaks of a *precipitate being formed,* which may be of any colour and consists of minute particles of the insoluble salt formed. These particles in the precipitate will sink to the bottom of the glass, which contains the fluid mixture, the rate of sedimentation or settling depending on the weight of the precipitate.

Suspension

If one wished to keep the precipitate or insoluble particles equally distributed throughout the fluid, one would require to add gum or such like substance to hold them in *suspension.* A *suspension,* then, is a mixture of a liquid and an insoluble substance, the latter being kept and held there uniformly distributed throughout the field by the addition of such a substance as gum.

Emulsification

The definition of suspension could equally well be substituted for *emulsification,* only the latter term has special reference to fats and oils.

If one were to pour oil into water, it would lie as a layer on the surface, but if one were to add it slowly, beating up the mixture at the same time, the fat globules would become broken up and dispersed evenly throughout the water forming an *emulsion.*

In milk the fat is emulsified, although when allowed to stand the fat, or cream, will settle in a layer at the top, but churn the milk and all the fat globules will adhere together to form a lump of fat called butter, and the fat emulsion cannot be reformed by simple shaking or mixing.

Ductibility or Malleability

These are terms applied to metals such as copper, silver, tin, etc.

Ductibility is the term used when substances can be drawn out into a thin wire, without breaking, and malleability is a term used to describe a metal that can be beaten out into a thin sheet.

PRACTICAL CHEMISTRY—CHAPTER 13

Chemical Terms

Expt. 65

Stable and Non-stable Substances

Show that sand, a stable substance, withstands the action of heat and chemicals such as strong sulphuric acid ; that sugar is quickly decomposed by these means and is said to be a moderately stable substance, whereas hydrogen peroxide on exposure to air is reduced to water and is therefore an example of an unstable substance.

Expt. 66

Deliquescence

Expose some calcium chloride to the air and note that it gradually liquefies.

Expt. 67

Efflorescent

Expose some crystalline sodium carbonate and note that it becomes white and powdery.

Expt. 68

Anhydrous Substance

Repeat Expt. 16 with the copper sulphate. The difference of the two powders is due to the loss of water of crystallization on heating.

Expt. 69

Oxidation

Heat a strand of copper wire, some copper foil, or a bright penny, over a flame. Note the formation of black copper oxide on the part exposed to the outside of the flame where oxygen is present.

Expt. 70

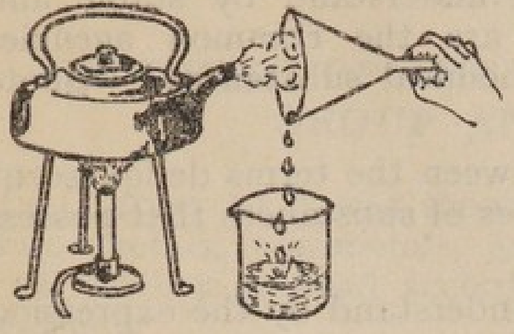
Reduction

Repeat previous experiment, plunging the copper foil when red hot into a little methyl alcohol contained in a test-tube.

The foil becomes bright at once due to the black copper oxide being reduced to metallic copper by the action of methyl alcohol which becomes oxidized.

EXPT. 71

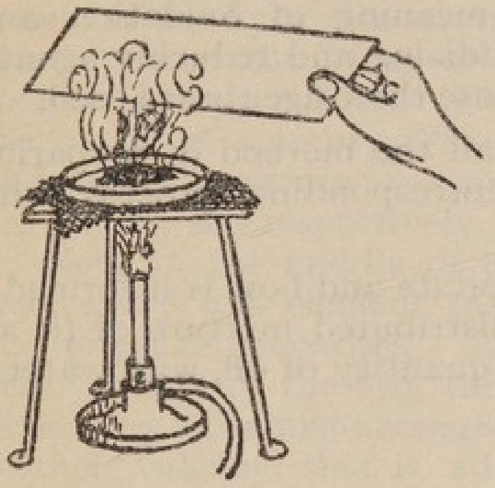
Distillation



Hold a funnel over the spout of a kettle containing boiling water; the steam condenses on hitting the cold funnel and drips from its lower edge as distilled water.

EXPT. 72

Sublimation



Hold a cold funnel or sheet of tin over a tin lid containing ammonium chloride.

Heat the lid and contents. Dense white smoke rises from the ammonium chloride and deposits as a white solid on the funnel or tin sheet. The white solid substance is sublimed ammonium chloride.

EXPT. 73

Volatile

Allow a drop of ether and a drop of water to fall on the table. The ether will dry up or vaporize rapidly, and for this reason is termed a volatile substance.

The water evaporates, but takes very much longer to do so.

EXPT. 74

Precipitation

Repeat Expt. 5.

A white insoluble powder sinks or is precipitated to the bottom of the tube.

EXPT. 75

Emulsification

Pour some olive oil on to some lime water in a test-tube. The oil remains on the surface and does not mix. Now shake the test tube, the oil mixes with the lime water and a thick emulsion results.

QUESTIONS

QUESTION 1.—What do you understand by stable and non-stable substances? What are the common agencies which cause decomposition of chemical substances?—Quote examples.

QUESTION 2.—Distinguish between the terms deliquescent and efflorescent, and give examples of substances that possess the particular property.

QUESTION 3.—What do you understand by the expression "the anhydrous form of a substance"? Illustrate your answer by describing in detail an experiment explaining how this form is prepared.

QUESTION 4.—What is the meaning of oxidation and reduction?—Give examples of oxidizing and reducing agents, and state for what particular purpose these agents are used.

QUESTION 5.—Describe in detail the method of preparing distilled water. What is the corresponding process with certain solids? Give an example.

QUESTION 6.—What is a precipitate and how is it formed? How can one obtain an evenly distributed mixture of (i) an insoluble powder, and (ii) a small quantity of oil, with water?

CHAPTER 14

GROUPING OF THE ELEMENTS—THE FIRST GROUP AND HYDROGEN

In Chapter 10 it was shown that elements can be classified into metals, non-metals, and metalloids, but they can be still further grouped according to their atomic weights and valency.

An atom of the element hydrogen is the lightest of all atoms and therefore is taken as 1, and all other elements are given an atomic weight which is fixed by comparing it with the weight of the hydrogen atom. The atom of oxygen is 16 times heavier than the atom of hydrogen; therefore the atomic weight of oxygen is 16. The atom of carbon is 12 times and the atom of mercury 200 times heavier than the atom of hydrogen; therefore the atomic weights of carbon and mercury are 12 and 200 respectively.

Now, in the middle of the last century it was discovered that every 8th element in the list had the same chemical and physical properties. The 1st element was similar to the 8th, the 2nd to the 9th, the 21st to the 28th, and so on. A little later a chemist arranged the elements in groups according to their valency, that is, all the elements with a valency of 1 were placed in one group, all those with a valency of 2 in a second, and so on until he had included all the known elements under 8 groups. It was found that this arrangement confirmed the accuracy of the grouping by atomic weight and proved that there was a definite relationship between the elements.

This natural arrangement of substances into groups, according to their atomic weights and valency, is known as the "Periodic Law," and can be best explained by studying the following table. Supposing the elements are arranged in rows of eight and in order of their atomic weights in this manner.

Valency							
Non.	Mono.	Di.	Tri.	Tetra.	Penta.	Di.	Mono.
He (4)	Li (7)	Be (9)	B (11)	C (12)	N (14)	O (16)	F (19)
Ne (20)	Na (23)	Mg (24)	Al (27)	Si (28)	P (31)	S (32)	Cl (35)
A (40)	K (39)	Ca (40)					

It will be seen that not only will the columns read vertically be of substances with similar character and properties, but will at the same time be arranged according to their valency. Helium, neon, and argon are similar substances and are non-valent; lithium, sodium, and potassium are also similar in character and are all monovalent; beryllium, magnesium, and calcium correspond closely and are all divalent; aluminium and boron agree and are trivalent, and so on.

As one progresses to substances of higher atomic weight, the classification becomes somewhat complicated. However, substances like chlorine, bromine, and sodium will be found to group together. Gaps will be found in the table, and these are gradually filled up as new substances are discovered; in fact, so accurate is the law that the nature and properties of some missing link has been fully described before its existence was known, and the description has agreed in every respect with the character of the substance when discovered.

The First Group and Hydrogen

The first group includes what are known as the "rare gases," *argon, helium, krypton, xenon,* and *neon,* and the common gas *hydrogen,* which although not actually belonging to this group is generally included in it. The former are present in the air in such small amounts and are of so little importance in practical chemistry that it is only sufficient to state that they exist.

They are non-valent elements, and therefore are never found in combination with any other element.

Hydrogen, however, is a gas of considerable importance and will be described in detail.

Hydrogen. Atomic weight 1.

History.—It was for a long time known as "inflammable air," and it was not until it was found to be one of the elements forming the compound water, that it was given the name "hydrogen," the name implying its connection with water.

Source in Nature.—In a free state it is found only in very small quantities, for example, emanating from volcanoes, but in combination with other elements it is extremely abundant, especially in the compound water.

Preparation.—There are many ways of preparing this gas, but the usual method is by the action of dilute sulphuric acid on zinc.

Equation: $Zn + H_2SO_4 = ZnSO_4 + H_2$

Other methods are by breaking up of the oxygen and hydrogen compound, water, by an electric current, or it can be produced in its purest form by the action of caustic soda (sodium hydroxide) on aluminium. The materials employed in this latter process are light and can be easily transported, and this is a very suitable method of generating hydrogen gas for army balloons in war time.

Properties.—Hydrogen is a colourless, odourless, tasteless gas and the lightest element known, having as already explained an atomic weight of 1. Being the lightest gas it (see Chapter 5) diffuses with greater rapidity than any other gas. It burns with a pale blue flame, giving off very little light but a great amount of heat. When it is mixed with oxygen or air in certain proportions and burnt, it explodes, forming the compound water. When subjected to great cold and pressure it changes from a gaseous to a liquid state. Hydrogen is not poisonous; so long as there is a sufficiency of oxygen no ill effects are felt, but animals deprived of the latter would soon die of suffocation if placed in a pure hydrogen atmosphere. It is very slightly soluble in water, but much more so in alcohol.

The oxides of hydrogen are *water* (H_2O) and *hydrogen peroxide* (H_2O_2).

Water (H_2O).

Pure water can only be obtained by distillation, by which means the solid matter and the gases dissolved in it are removed.

Water as it occurs naturally is really a dilute solution of many salts and gases, and its flavour will depend mainly on the particular salt that is dissolved. Some water contains large amounts of salts and are of great value for their medicinal action. Such waters are called " natural mineral waters " and are found in deep springs, for example, those at Epsom, Bath, Vichy, etc.

Waters that come from a deep subterranean source are generally hard, and although they are the most pleasant to drink they may be undesirable—for example, in dispensing the presence of a dissolved salt may alter the composition of a medicine. The hardness is due to the presence of calcium and magnesium carbonate and calcium and magnesium sulphate ; when the former are present in the greatest amount the water is from a chalk strata and is said to possess temporary hardness because on boiling this hardness is removed. Whatever hardness remains after boiling is termed permanent hardness, and is due to the sulphates of magnesium and calcium.

The term " hard " applied to water has originated from the fact that it is very difficult to form a lather with soap with such waters, whereas with a soft water, that is, one without these dissolved salts, a lather forms almost immediately. Advantage has been taken of this property to determine the degree of hardness of a water, and the test commonly in use is to add small amounts of soap solution to a measured quantity of water, shaking between each addition and noting the amount of soap solution that is required to form a permanent lather. This would give the total hardness, and to find out how much is temporary and how much permanent the water must be boiled to remove the former and whatever hardness remains is permanent hardness. The temporary hardness is calculated by subtracting the permanent hardness from the total.

Total hardness — temporary hardness = permanent hardness.
(by boiling)
Total hardness — permanent hardness = temporary hardness.

A hard water, although sometimes the best water for drinking, is not always suitable for factories and laundries. In the first place the presence of the salt may interfere with the chemical process carried out or may cause furring, that is, a deposit of chalk in boilers leading to a waste of fuel and to a grave risk of the boiler bursting.

In the second instance the amount of soap required to produce a lather for the proper washing of clothes is so great that to run a laundry with a hard water would be a very expensive concern.

A hard water may be the only water supply, in which case it must be softened if it is to be used for the above purposes. This is usually done by what is called the " lime and soda " method ; the lime causes the calcium and magnesium carbonate to be precipitated and this removes the temporary hardness, while the washing soda (sodium carbonate) by chemical action alters the magnesium and calcium sulphate to magnesium and calcium carbonate, which are both insoluble and are precipitated. The solid matter resulting from the treatment is then removed by filtration or sedimentation.

HYDROGEN PEROXIDE

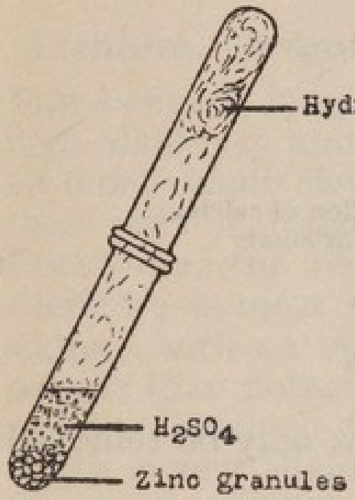
Is a liquid which, similar to water in appearance, is very easily decomposed, giving off oxygen and being reduced eventually to water. It is used extensively (as an antiseptic) in surgery, its action being due to the oxygen liberated, and for the same reason it is employed as a bleaching agent.

PRACTICAL CHEMISTRY—CHAPTER 14

EXPT. 76

Hydrogen

Drop a few granules of zinc into some dilute sulphuric acid contained in a test-tube.



Collect the gas evolved in a dry test-tube (see diagram). Hydrogen being a very light gas rises and displaces the air from the tube.

Remove and apply a lighted match to the mouth of the tube. A small explosion occurs with the formation of moisture on the inner surface of the tube, two atoms of hydrogen from the acid having combined with one atom of oxygen from the atmosphere to form water.

EXPT. 77

Water

Hardness—Permanent

Shake together some distilled water and a small piece of soap in a test-tube. A smooth foamy lather immediately forms, and if shaking is continued the lather will fill the whole of the tube, indicating a very soft water.

Add a few crystals of magnesium or calcium sulphate to the tube, the lather will disappear and will not reappear on repeated shaking or boiling, showing that it has now become a very hard water.

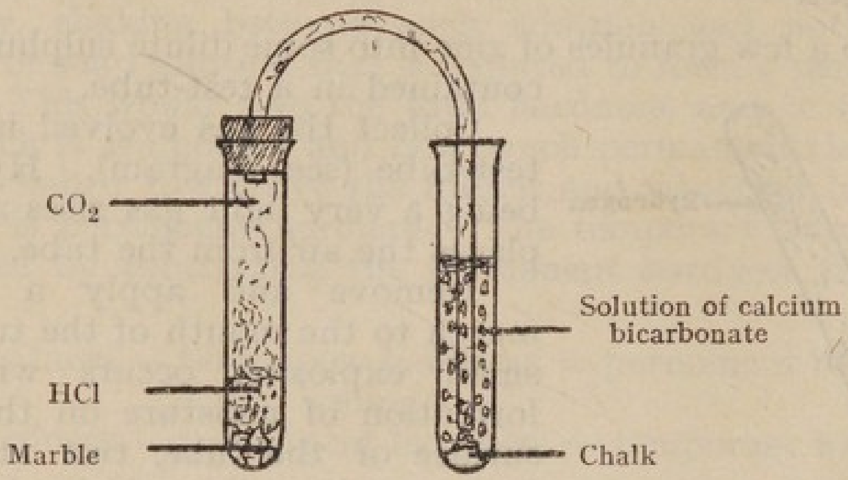
Now add a few crystals of washing soda to the tube and shake, the lather will immediately form, the magnesium or calcium sulphate has become converted to insoluble carbonate which does not prevent the formation of a lather.

EXPT. 78

Temporary

Add some 50 per cent. hydrochloric acid to a carbonate (marble chips or crystallized sodium carbonate) contained in a test-tube fitted with a tube as shown in diagram. Bubble the carbon dioxide gas evolved through distilled water containing a small pinch of chalk in a test-tube, when some of the chalk will dissolve. Remove the tube and allow the undissolved chalk to settle.

Pour off the clear solution into another tube. This water now contains temporary hardness. To half of this water add a small piece of soap and shake—no lather will form; boil for a few minutes to drive off the carbon dioxide and shake, a lather will form, showing that temporary hardness may be removed by boiling.



To the remaining half add a small pinch of lime and shake, then add a small piece of soap and shake, a lather will form, showing that temporary hardness may also be removed by the addition of lime.

Expt. 79

Hydrogen Peroxide

Pour a few drops of hydrogen peroxide on some moist earth—frothing due to the liberation of the oxygen occurs.

Heat a little in a test-tube to drive off the oxygen and repeat—frothing does not now occur, showing that the hydrogen peroxide has decomposed to water.

QUESTIONS

Question 1.—Explain briefly what you know of the " Periodic Law."

Question 2.—How is hydrogen prepared and what are its properties ?

Question 3.—Discuss the features of hard and soft waters. How can a hard water be made soft ?

CHAPTER 15

THE SECOND GROUP—MONOVALENT ELEMENTS —THE ALKALI METALS

Lithium, Sodium, and Potassium, also Ammonium

The elements included in this second group are noted for their high alkalinity and for the solubility of their salts. They are all monovalent elements, and the alkalinity and solubility of their salts decrease with the decrease in their atomic weights. Potassium, the heaviest, has an atomic weight of 39, and therefore is most strongly alkaline and most soluble; while sodium with an atomic weight of 23 is less alkaline and less soluble than potassium, but more so than lithium, the atomic weight of which is 7 and the alkalinity and solubility of which is very slight.

Lithium (Li). Atomic weight 7

Source in Nature.—Lithium is found in combination with a few rare minerals.

Preparation.—It is prepared by the action of an electric current on the chloride of lithium.

Properties.—It is a silver-white metal softer than lead and can be cut easily with a knife, and is the lightest known solid.

Compounds—

Lithium Oxide (Li_2O) is formed when the metal burns in the presence of air. The oxide is soluble in water and when dissolved forms lithium hydroxide.

Lithium Carbonate (Li_2CO_3) is precipitated when some sodium or potassium carbonate is added to a solution of the nitrate or chloride of lithium. It is only slightly soluble.

When compounds of lithium are burned in a bunsen flame they produce a red coloration of the flame.

Lithium salts are used medicinally in the treatment of gout, rheumatism, and stone in the kidney.

Sodium (Na). Atomic weight 23

Source in Nature.—The commonest and most plentiful source of sodium is from common salt, sodium chloride (NaCl), which is found in enormous quantities as rock salt in various countries in Europe and as a solution in sea water.

Sodium in the form of sodium nitrate (saltpetre) is found abundantly in Chili.

Preparation.—Sodium is best prepared by the action of electricity on caustic soda (NaOH) or sodium chloride.

Properties.—It is a soft white metal forming the oxide when exposed to the air, to prevent which, it must be kept stored in naphtha or such like substance. It decomposes water with the formation of hydrogen.

Compounds—

Oxides.—Two compounds with oxygen are formed, *sodium monoxide* (Na_2O) and *sodium peroxide* (Na_2O_2).

SODIUM HYDROXIDE (NaOH). Caustic soda

Source in Nature.—Caustic soda is far too active a substance to be found pure in nature. This applies also to caustic potash.

Preparation.—The simplest method of preparing caustic soda is by dissolving the pure metal sodium in water and finally evaporating the excess water.

On a large scale it is produced by the action of lime on a boiling solution of sodium carbonate (washing soda).

$$Na_2CO_3 + Ca(OH)_2 = 2Na(OH) + CaCO_3$$

The caustic liquor is then filtered from the chalk and evaporated until it solidifies on cooling. When this stage is reached it is poured into pans, allowed to set, broken up and finally packed into drums.

Caustic soda can also be manufactured by the electrolysis of brine (sodium chloride solution).

Properties.—It is a white, powerfully caustic and highly deliquescent solid which is very soluble in water, much heat being given out while it dissolves. It takes up CO_2 readily from the air, forming sodium carbonate.

Uses.—It is used in the manufacture of soap and for cleaning purposes in dye works.

SODIUM CHLORIDE (NaCl). Common salt

Source in Nature.—Sodium chloride is obtained by the evaporation of sea water or from rock-salt mines. It usually contains impurities, such as magnesium and other salts dissolved in the sea water.

Preparation.—Pure sodium chloride is prepared by adding hydrochloric acid to a saturated solution of common salt. Sodium chloride is deposited while the impurities, that is, the other salts, are kept in solution.

Properties.—Sodium chloride is a soluble substance, one part dissolving in three parts of cold water.

SODIUM CARBONATE ($Na_2CO_310H_2O$). Washing soda

Source in Nature.—Sodium carbonate was formally obtained from the ash left on burning seaweed.

Preparation.—It may be manufactured by one of the following processes :—

The Leblanc Process consists of two stages. First, the production of sodium sulphate and hydrochloric acid by the action of sulphuric acid, and then by the conversion of the sodium sulphate into sodium carbonate by heating with chalk and coal.

The Ammonia-Soda Process.—In this process ammonia and carbon dioxide gases are pumped into a strong solution of salt, and sodium bicarbonate is obtained as a precipitate. The sodium bicarbonate is heated to drive off the CO_2, and is thus converted into sodium carbonate.

The Modern Electrolytic Method.—Brine is electrolysed in the presence of carbon dioxide gas into sodium carbonate.

In the Leblanc process, as a result of using commercial sulphuric acid, arsenic may be present as an impurity in sodium carbonate. Similarly, ammonia may be present in the soda made by the ammonia-soda process.

Properties.—When sodium carbonate, which is a crystalline substance, is exposed to the air it gives up its water of crystallization and becomes a white amorphous powder ($Na_2CO_3H_2O$).

Sodium carbonate is very soluble in water and strongly alkaline in character.

Uses.—It is used to soften water.

SODIUM BICARBONATE ($NaHCO_3$). Baking soda

Source in Nature.—Bicarbonate of soda does not occur in nature.

Preparation.—It can be prepared as described in the ammonia-soda process, or by the action of carbon dioxide on sodium carbonate.

Properties.—Its alkalinity and solubility is much less than that of the carbonate. When strongly heated it loses carbon dioxide and becomes sodium carbonate.

SODIUM SULPHATE (Na_2SO_4), Glauber's Salt; and SODIUM NITRATE ($NaNO_3$). Saltpetre

Glauber's salt is used as an aperient in medicine, while sodium nitrate, known as Chili Saltpetre, is used as an artificial manure and in the preparation of nitric acid.

When salts of sodium are heated in a flame they colour the latter an intense yellow colour.

PRACTICAL CHEMISTRY—CHAPTER 15

Sodium—Potassium—Ammonium

Expt. 80

Caustic Soda and Caustic Potash

Demonstrate the great solubility of these two substances by mixing in a test-tube equal quantities of one or other of them with water. Note the generation of heat during the process.

Test with red litmus paper, which turns blue.

(*Caution.*—Do not touch the caustic with the fingers as it rapidly dissolves the skin.)

Leave a little on a glass on the bench overnight to show deliquescence.

Expt. 81

Sodium Chloride

Examine with a lens crystals of salt, and note their cubical shape.

Add one part of salt to two parts of water in a test-tube, and shake so as to dissolve as much salt as possible.

Hold the tube vertically and mark off the level of the undissolved salt. Now boil the solution and compare the height of the undissolved salt; it will be almost the same, showing that the solubility of the salt is not affected by heat.

Dip a red and blue litmus paper in the solution; the colour is unchanged, for sodium chloride is a neutral salt.

Expt. 82

Sodium and Potassium Carbonate

Repeat Expt. 67, showing that both these substances effloresce.

Demonstrate that they are readily soluble in cold and hot water by adding one part of the salt to two parts of water.

Dip a red litmus paper in the solution; it turns blue, indicating an alkaline salt.

Expt. 83

Sodium—Potassium Bicarbonate

Mix one part of either of the bicarbonates with five parts of water and shake. They are fairly soluble. Dip a piece of red litmus paper in the solution; it turns blue.

Now gently heat the solution and shake; its solubility does not increase, but when it is raised to boiling point, bubbles of CO_2 escape and all the salt will dissolve.

The heat has driven off one molecule of CO_2 from the bicarbonate, which has become the carbonate, which is very soluble.

EXPT. 84

Sodium Sulphate

Repeat Expt. 67, with Glauber's salt; it effloresces.

Heat a few crystals in a tin lid; the crystals become liquid, having dissolved in their own water of crystallization. It is therefore a very soluble substance.

Shake equal quantities of the salt and water in a tube, touch the tube; it has become intensely cold, showing that the substance in dissolving absorbs heat.

Test the reaction of the solution; it is neutral.

CHAPTER 15 (*continued*)

Potassium (K). Atomic weight 39

Source in Nature.—This element occurs widely distributed throughout the world in combination with other elements. It occurs in granite as a compound of silica and aluminium, and in the ash of plants as potassium carbonate.

It is found in sea water and some mineral springs, but its chief source is from saltpetre—potassium nitrate—from Chile and potassium chloride in salt mines in Germany.

Preparation.—The best method of producing potassium in the pure state is by the decomposition or breaking up of caustic potash—potassium hydroxide—by electrolysis.

Properties.—Potassium is a shining white, very soft metal and resembles sodium in many respects.

It is easily oxidized on exposure to air, and thus has to be protected by storage in naphtha. It liberates hydrogen from water and forms caustic potash.

COMPOUNDS OF POTASSIUM

Compounds—

Potassium oxide (K_2O) is a grey brittle substance, and potassium peroxide a yellow powder formed on heating potassium in air.

POTASSIUM HYDROXIDE (KOH). Caustic soda

Preparation.—It is prepared in the same manner as sodium hydroxide, only potassium carbonate replaces sodium carbonate.

$$K_2CO_3 + Ca(OH)_2 = 2KOH + CaCO_3$$

In the electrolytic method, potassium chloride takes the place of common salt.

Properties.—Caustic potash is a white brittle solid, very deliquescent, dissolving readily in water with the evolution of great heat. It is the most powerful alkali in common use. It is purified by alcohol, in which it is soluble.

POTASSIUM CHLORIDE (KCl)

This salt is mined in Germany and is chiefly used for the production of other potassium salts.

POTASSIUM CHLORATE ($KClO_3$)

This salt is prepared by electrolytic action on a saturated solution of potassium chloride.

It is used commercially in the manufacture of matches and fireworks, as it readily parts with its oxygen.

Medicinally, it is used as a throat antiseptic.

POTASSIUM IODIDE (KI)

Preparation.—Potassium iodide is prepared in a somewhat complicated manner from the reaction of iodine and a solution of potash.

Properties.—The crystals of potassium iodide are easily recognized by their distinct cubical shape. This salt is very soluble in water and is used extensively as a medicinal drug, in photography and in chemistry generally.

POTASSIUM CARBONATE (K_2CO_3)

Preparation.—The carbonate was originally obtained from wood ashes, the salt being dissolved out by the addition of water. There is another method of preparation in which magnesium carbonate, potassium chloride, and carbon dioxide are the chemicals used.

Properties.—Potassium carbonate is deliquescent and very soluble in water, the solution obtained being strongly alkaline.

POTASSIUM BICARBONATE

Prepared in the same way as soda bicarbonate, potassium carbonate being used instead of the sodium salt.

POTASSIUM SULPHATE (K_2SO_4)

This salt is found in combination with magnesium, and is used extensively as a manure.

POTASSIUM NITRATE (KNO_3). Nitre or Saltpetre

Source in Nature.—Nitre is found in India and South America, especially in places where there is little rainfall. It appears on the surface of the ground as a white crust.

It is extracted with water and the solution evaporates when crystals, which are later purified, are obtained.

Properties.—Potassium nitrate crystallizes out from solution into long crystals, which are not deliquescent, and although soluble in water they are insoluble in alcohol.

It is used in the preparation of nitric acid and in the manufacture of gunpowder.

When heated in a flame the salts of potassium produce a violet colour.

Ammonia (NH_3)

Certain groups of atoms of different elements, when acted on by other substances, do not break up, but retain their combination—in this manner, behaving as though they were elementary bodies.

These groups of atoms are known as *compound radicals.* Ammonia is not an element, it is the best example of a compound radical. If acted on by hydrochloric acid or other acids it does not break up into its component parts, nitrogen and hydrogen, but tacks itself on unchanged to the acid group, and forms an ammonium salt, ammonium chloride.

$$NH_3 + HCl = NH_4Cl$$

Ammonia has no right therefore to be included in the second group of elements, but it is usually placed here, because its salts resemble the salts of sodium and potassium in a very marked degree.

Source in Nature.—Ammonia occurs in manure and rubbish heaps.

Preparation.—It is prepared from the "ammoniacal liquor," a by-product of gas works, by treatment with lime.

Properties.—A light colourless gas with a pungent odour and very soluble in water forming an alkaline solution.

It is used for refrigerating purposes.

Compounds of Ammonia:

AMMONIUM CHLORIDE (NH_4Cl). Sal ammoniac

Source in Nature.—It occurs in the vicinity of volcanoes.

Preparation.—Ammonium chloride is formed when hydrochloric acid gas and ammonia combine together.

Properties.—It occurs as crystals or as a sublimate, being known as "sal ammoniac." It is very soluble in water and absorbs a great deal of heat on dissolving.

It is also used as a soldering flux and in solution as one of the chemicals of a Leclanché cell.

AMMONIUM CARBONATE ($(NH_4)_2CO_3$). Smelling salts

Preparation.—Commercial ammonium carbonate is manufactured by heating a mixture of chalk and ammonium chloride. This commercial carbonate contains a certain amount of the bicarbonate and for laboratory purposes where the pure salt is required, this impurity is converted into the carbonate by the addition of ammonia gas to a solution of the commercial salt.

Properties.—Ammonium carbonate is easily decomposed by heat into ammonia and carbonic acid gas. It is insoluble in alcohol, but the carbamate is soluble and forms "sal volatile." It is used for medicinal purposes.

AMMONIUM SULPHATE ($(NH_4)_2SO_4$)

Source in Nature.—Ammonium sulphate is found in the neighbourhood of volcanoes.

Preparation.—It is prepared by neutralizing ammonium hydroxide with sulphuric acid and allowing it to crystallize out.

Properties.—It is soluble in water although insoluble in alcohol.

Ammonium sulphate is employed for manurial purposes.

AMMONIUM NITRATE (NH_4NO_3)

Preparation.—Ammonium nitrate is formed when ammonium hydroxide is neutralized by nitric acid.

Properties.—It is a crystalline substance, being deliquescent and very soluble in water. It decomposes on being heated, forming nitrous oxide gas and water. Ammonium nitrate is largely used for making explosives and for the manufacture of nitrous oxide gas.

PRACTICAL CHEMISTRY—CHAPTER 15 (*continued*)

EXPT. 85

Potassium Chlorate

(*a*) Shake together in a test-tube one part of potassium chlorate and three parts of water.

It is sparingly soluble. Heat the contents of the tube and the whole will dissolve showing that it is more soluble in hot water.

(*b*) Place a few crystals in a mortar and rub lightly, the crystals explode with a sharp crackling noise. Heat a little in a test-tube, it melts and commences to bubble. Hold a glowing match in the tube, it burns more brightly, due to the freed oxygen.

EXPT. 86

Sodium and Potassium Flame

Sprinkle a pinch of common salt on to a bunsen flame and note the yellow colour imparted to the flame. Repeat, using a potassium salt and note the lilac colour.

EXPT. 87

Ammonium Hydroxide

Add some strong ammonia solution to water and note the smooth soapy feel of the solution.

Test with red litmus paper to show strong alkalinity.

EXPT. 88

Ammonium Chloride

Repeat Expt. 72.

Shake together one part of the salt with three parts of water in a test-tube, it rapidly dissolves becoming cold on doing so. Test the reaction; it is neutral.

EXPT. 89

Ammonium Carbonate

Note the strong smell of ammonia.

Dissolve one part of the salt in four parts of water and note the soapy feel imparted to the water. Test its reaction, it is alkaline.

Expose it to the air; it becomes powdery, it effloresces.

QUESTIONS

QUESTION 1.—What is the method of preparation and the special properties of sodium hydroxide, sodium carbonate, and sodium sulphate, and for what purposes are they used?

QUESTION 2.—What is the source and state the properties of potassium? Why is it called an alkali metal?

QUESTION 3.—Enumerate the particular properties possessed by and uses of potassium hydroxide, potassium chlorate, and potassium nitrate.

QUESTION 4.—What is meant by the term compound radical? Mention an example and show by an equation what occurs when it is acted on by an acid.

QUESTION 5.—How are ammonium chloride and ammonium carbonate prepared and what are their respective properties? By what commercial names are they known?

CHAPTER 16

THE THIRD GROUP—MONOVALENT AND DIVALENT ELEMENTS

Copper—Silver—Gold—Mercury—Platinum

THE elements named above bear a great resemblance one to the other, both in their chemical and in their physical properties, and therefore are dealt with in one group.

They are all capable of forming either monovalent or divalent compounds.

Copper (Cu). Atomic weight 63·5

Source in Nature.—Copper in almost a pure state may be found as veins running through certain rocks. More often it occurs as a compound of which the commoner are red copper oxide (Cu_2O), and copper sulphide (copper pyrites), which always contain iron as an impurity, and also as copper carbonate.

Preparation.—Copper is obtained from its ores by roasting the latter, a special process being necessary to remove the iron impurities from the sulphide.

Properties.—Copper is a reddish-brown lustrous metal which is extremely malleable, and this allows it to be beaten into thin sheets, and ductile, therefore capable of being drawn out into wires which are used extensively in electrical fittings, and as it is an excellent conductor of electricity it is especially suitable for the latter use.

The air has a slow action on it, but neither hydrochloric nor sulphuric acid if cold dissolve it.

Nitric acid, however, both dilute and concentrated, has a very vigorous action on this metal. Copper is used in the manufacture of alloys, of which brass and bronze are the more important.

CUPRIC SULPHATE ($CuSO_4(5H_2O)$). Blue Stone. Blue Vitriol

Source in Nature.—Copper sulphate may occur in pure crystalline form in copper mines where sulphur is also present.

Preparation.—It can be prepared by the action of hot sulphuric acid on copper.

Properties.—The crystals are blue in colour, but when heated they lose their water of crystallization and become a white amorphous mass.

Copper sulphate is very soluble in water and is " acid," turning blue litmus red.

Copper sulphate is used in the dye industry, and as a caustic in the treatment of ulcers. It is also useful in the cleansing of ponds, water supplies, and baths, killing fungi and vegetable growths which foul the water, and is effective in destroying certain insect pests which attack wheat. Fehling's, Benedict's, and Bang's solutions are prepared from copper sulphate, and are used for the estimation of sugar in substances.

CUPRIC ARSENITE. Scheeles Green

Under the name of " Paris Green " is used extensively, mixed with road dust, as a poison to destroy mosquito larvæ in ponds.

VERDIGRIS

When acetic acid acts on copper, a greenish-blue substance —copper acetate or " verdigris "—is formed which on exposure to air is slowly converted into the carbonate of copper.

All copper salts burn with a greeny-blue flame.

Silver (Ag). Atomic weight 108

Source in Nature.—Silver (in the pure state) occurs in many countries, especially in South America, and also as the sulphide of silver.

On account of its comparative scarcity, it is included amongst the rare metals, and is utilized in most countries for coinage.

Pure silver in a colloidal state is used medicinally. The silver in very minute microscopic particles is suspended in water, the mixture being known as colloidal silver.

Preparation.—Silver is obtained from its ores by smelting.

Properties.—Silver is a shining white metal. It is the best known conductor of heat and electricity, its cost alone prohibiting its use for wiring purposes.

On account of its malleability it can be fashioned into any shape or form.

It is not oxidized in the presence of air, but is rapidly coated with a black tarnish when exposed to sulphur gas.

As already stated, it is used in the manufacture of coins, but requires the addition of a small amount of copper to harden it and prevent it wearing away too quickly. Hydrochloric acid acts slightly on silver, sulphuric acid only when boiling, but nitric dissolves it readily.

Many so-called silver articles are merely coated with silver, and are known as electro-plate, the process by which the coating is effected being that of electrolysis.

Compounds of Silver

SILVER BROMIDE ($AgBr$)

Source in Nature.—Silver bromide occurs in nature in combination with two minerals.

Preparation.—It can be prepared by mixing together ammonium bromide and silver nitrate.

Properties.—It has a yellowish colour and is extremely sensitive to light, being much used in photography.

It is insoluble in water.

SILVER CHLORIDE ($AgCl$) and SILVER SULPHIDE (Ag_2S)

Are salts both of which are found in nature.

SILVER NITRATE ($AgNO_3$). Lunar Caustic

Preparation.—It is readily prepared by the action of nitric acid on pure silver.

Properties.—Silver nitrate occurs in square colourless crystals which are anhydrous, non-deliquescent, and very soluble in water.

It readily melts, and when made into stick form is used by the surgeon as a caustic.

Pure silver nitrate is not affected by light, but if organic matter is present it blackens, and for this reason it is invariably kept in a coloured glass jar.

Silver forms compounds with protein, two of the more important ones being *Argyrol* and *protargol*, which are both used in the treatment of eye and other conditions.

Gold (Au). Atomic weight 197

Source in Nature.—Gold is a rare and therefore precious metal, and is found always in the pure state either as veins in the solid rock or as small nuggets in sand deposits. It is widely distributed in small quantities throughout the world.

Preparation.—It is dissolved out from the crushed rocky mass by sodium cyanide or extracted by adding mercury. Purification of the metal is usually effected by electrolytic means.

Properties.—Gold is a heavy yellow metal. It is a very good conductor of heat and electricity, and like silver is extremely malleable and ductile. It is unaffected by air, and is therefore a favourite substance for ornaments. It resists the action of most acids, but is dissolved by "*aqua regia,*" a mixture of nitric and hydrochloric acids, forming auric chloride. It is too soft in the pure state for coins, and therefore is hardened by the addition of copper. A colloidal form of pure metallic gold similar to colloidal silver is used sometimes in medical treatment.

Compounds of Gold

The oxides are three in number, but are of no importance.

GOLD CHLORIDES

There are three chlorides of gold, the only important one being auric chloride ($AuCl_3$), which as we have seen is prepared by dissolving gold in aqua regia, a mixture of nitric acid and hydrochloric acid, and then evaporating the liquor. The excess hydrochloric acid is driven off by gently heating.

Properties.—The crystals of auric chloride are reddish-brown, deliquescent, and very soluble in water. Auric chloride is much used in photography.

CHAPTER 16 (*continued*)

Mercury (Hg). Atomic weight 200. " Quicksilver "

Source in Nature.—Mercury is found in the metallic state, and also in combination with other elements. In combined form it is known as cinnabar, the sulphide.

Preparation.—Mercury in a partially purified state is extracted from its ores by smelting and distilling in a furnace. The impurities, for example dirt, zinc, lead, etc., are removed by dissolving them out, either with water or if insoluble with nitric acid.

Properties.—Mercury is the only metal which is liquid at ordinary temperatures. It remains liquid until cooled to 38·8° C. below zero, and its boiling point is high, 357° C. It expands and contracts evenly and quickly when heated or cooled, and these qualities combined with its property of not melting or sticking to glass and its heaviness make it a most suitable substance for the manufacture of barometers and thermometers. Mercury is silver-white in colour, and is very mobile, and for these reasons has been named " quicksilver."

In pure air mercury does not tarnish, but if sulphur is present it does so like silver and copper. It resists the action of hydrochloric acid, but it is dissolved by nitric and by strong sulphuric acid. Mercury even at ordinary temperatures freely combines with most metals, forming alloys which are specially named " amalgams." Copper, silver, gold, zinc, and tin are those metals most readily affected.

Amalgams of gold and tin are used in dentistry.

Compounds of Mercury

OXIDES OF MERCURY

These are two in number, namely mercurous oxide (Hg_2O) and mercuric oxide (HgO).

MERCUROUS OXIDE—Is a very unstable compound, being readily decomposed into mercuric oxide and mercury.

$$Hg_2O = HgO + Hg$$

MERCURIC OXIDE (HgO). Red oxide of Mercury

The red oxide is produced when mercury is subjected to prolonged heating in air at a temperature of about 300° C. A more convenient way, however, is by heating mercuric nitrate.

Mercurous Salts

MERCUROUS CHLORIDE ($HgCl$). Calomel

Preparation.—It may be prepared by heating together in the correct proportions mercuric chloride and mercury, the calomel condensing on cooling.

$$HgCl_2 + Hg = 2HgCl$$

Another and more common method is by heating mercury, mercuric sulphate, and common salt, the calomel condensing on cooling.

$$HgSO_4 + Hg + 2NaCl = 2HgCl + Na_2SO_4$$

Properties.—Mercurous chloride is a white solid supplied either in masses or as an amorphous powder. Calomel is insoluble in water and weak acids, but dissolves in boiling strong nitric acid. It also decomposes when boiled with strong hydrochloric acid.

Mercuric Salts

MERCURIC CHLORIDE ($HgCl_2$). Corrosive sublimate

Preparation.—Corrosive sublimate is prepared by heating mercuric sulphate and common salt together.

$$HgSO_4 + 2NaCl = HgCl_2 + Na_2SO_4$$

The mercuric chloride vaporizes and condenses on cooling.

Properties.—It occurs as long needle-shaped crystals, moderately soluble in cold and very soluble in hot water, in alcohol and ether. When a hot solution cools, the long crystals are deposited. Corrosive sublimate combines readily with ammonia compounds.

Unlike calomel it is a very violent poison, and one of the most powerful antiseptics. It coagulates albumen, and therefore has little power of penetration, and for this reason is not altogether suitable for disinfecting purposes.

MERCURIC IODIDE (HgI_2)

It is formed by heating mercury and iodine together, or by the action of potassium iodide on mercuric chloride when it is precipitated.

Properties.—Mercuric iodide is a brilliant red powder which changes its colour to yellow on gently heating. Mercuric iodide is insoluble in water, but soluble in ether, hot alcohol, and potassium iodide solution.

It is used extensively as an antiseptic in surgery.

Palladium (Pd). Atomic weight 106·7

This metal is found associated with gold and platinum in certain mineral ores.

Platinum (Pt). Atomic weight 195

Source in Nature.—It occurs only in the metallic state in ores which also contain gold and palladium.

Properties.—It is a white metal, very heavy and extremely difficult to melt.

Platinum is not acted on by air at any temperature. This makes it of value for ornaments and for dishes in which substances are heated to a high temperature. Almost any substance can be heated in a platinum dish without the latter corroding, one of the exceptions being aqua regia.

It is a soft metal and often requires hardening by introducing a little of the metal iridium. Being very malleable and ductile, it can be beaten out into a thin foil, or can be drawn out into a fine wire.

PRACTICAL CHEMISTRY—CHAPTER 16

Third Group

Expt. 90

Copper

Heat a piece of copper in a flame, and note the green colour of the flame which is characteristic of copper and its compounds.

Examine a piece of copper wire; it has a reddish-brown lustrous appearance. Bend the wire with the fingers, it bends without cracking. It can be flattened out.

Hold one end of the wire in a flame; heat travels quickly along the wire, showing that it is a good conductor of heat. In Expt. 55 copper wire was used to conduct electricity between the two plates, it is therefore a conductor of electricity.

Expt. 91

Cupric Sulphate

Repeat Expt. 16.

Examine crystals. Shake together one part of the salt with four parts of water in a test-tube; the salt dissolves, forming a blue solution. Dip a piece of blue litmus paper in the solution; it turns red.

Place a clean iron nail in the solution, for a few minutes, remove and note the pure copper deposited on the nail.

Add a few drops of ammonia to the solution; it goes a deep blue colour. This is a test for copper compounds.

Expt. 92

Silver

Examine a silver coin; it is hard, because it is alloyed with a small proportion of copper.

Hold the edge of the coin in a flame. Heat travels quickly across the coin, therefore it is a good conductor of heat.

Expt. 93

Silver Nitrate

Examine a stick of silver nitrate; it is rolled in black light-proof paper to prevent decomposition on exposure to light.

(*Caution.—Be careful not to let the stick touch the fingers, as it stains.*

Dissolve a small piece in distilled water in a test-tube. It rapidly dissolves forming a clear colourless solution. Add a little of this solution to some tap water in another test-tube. Should it contain any salt, it will become hazy due to the formation of insoluble silver chloride.

Expt. 94

Mercury

Note its colour, weight, and mobility.

Repeat Expt. 21. (If available mercury from broken thermometers will do.) Heat a little in a test-tube; it soon begins to boil. The vapour will deposit on the upper cool part of the tube forming a mirror. This shows that mercury can be sublimed.

Expt. 95

Mercurous Chloride—Calomel

Rub a little of the white powder between the fingers. It has a smooth feel like flour.

Shake up a pinch in half a test-tube of water. It is insoluble.

Expt. 96

Mercuric Chloride—Corrosive Sublimate

Rub a pinch of the powder between the fingers. It has a harsh granular feel. Note its weight.

Add one part to two parts of water and shake; it is sparingly soluble. Now heat it, and its solubility rapidly increases. Cool and note the long, white, silky needle-like crystals which form.

(CAUTION.—*This is a dangerous poison—wash your hands well.*)

EXPT. 97

Mercuric Iodide

Examine a little of the powder—note the brilliant red colour. Heat a pinch in a test-tube; it turns yellow and remains yellow on cooling.

Heat a pinch in water; it is insoluble. Add a little to some alcohol and shake; it immediately dissolves.

QUESTIONS

QUESTION 1.—What is the natural source and describe the properties of copper and any one of its salts you like to name?

QUESTION 2.—Mention two salts of silver and describe the source, properties, and uses of one of them.

QUESTION 3.—How is mercury prepared from its ores? Describe its properties and its uses.

QUESTION 4.—Distinguish between mercurous and mercuric chloride.

CHAPTER 17

THE FOURTH GROUP—DIVALENT ELEMENTS

Magnesium—Calcium—Strontium—Barium—Cadmium—Zinc, and Radium

MAGNESIUM, calcium, strontium, and barium are known as the alkaline earth metals, and most of their salts are alkaline in reaction.

Magnesium (Mg). Atomic weight 24

Source in Nature.—Many of the rocks forming the earth's crust are composed of magnesium compounds, of which the commoner are magnesium carbonate, magnesium chloride, magnesium sulphate, and magnesium silicate.

Preparation.—The pure metal magnesium is obtained by the electrolysis of the crude magnesium chloride.

Properties.—It is a silver-white metal which is malleable and ductile, and is used beaten out in the form of a ribbon or drawn out into a wire. The metal slowly tarnishes on exposure to air, forming magnesium oxide. It burns readily in air and oxygen, giving off a dazzling white light, and is much used in photography and fireworks. Because of its lightness it is sometimes alloyed with aluminium. Dilute solutions of acids easily dissolve magnesium, forming magnesium salts.

Compounds

MAGNESIUM OXIDE (MgO). Magnesia

Preparation.—This is prepared by burning magnesium in air or oxygen.

It may also be derived from magnesium carbonate or magnesium nitrate by heating strongly.

$$MgCO_3 + \text{Heat} = MgO + CO_2$$

Magnesia is a white powder which withstands great heat before it melts. It is for this reason used in the manufacture of fire bricks. It is slightly soluble in water forming the hydroxide, which gradually is converted into magnesium carbonate by absorbing the carbon dioxide from the air.

$$Mg(OH)_2 + CO_2 = MgCO_3 + H_2O$$

Magnesium hydroxide can also be formed by the interaction of magnesium chloride and caustic soda.

$$MgCl_2 + 2NaOH = Mg(OH)_2 + NaCl$$

MAGNESIUM CARBONATE ($MgCO_3$)

Source in Nature.—It occurs in nature as the mineral magnesite or dolomite. There are two carbonate compounds, the light magnesium carbonate and the heavy magnesium carbonate.

Preparation.—They are both prepared by the interaction of solutions of magnesium sulphate and sodium carbonate, the only difference being that in the case of the heavy carbonate the solutions are mixed when boiling, whereas in the case of the light carbonate they are mixed cold and then boiled.

Light and heavy magnesia results from the roasting of their respective carbonates.

Properties.—Magnesium carbonate is alkaline, insoluble in water, but is acted on by acids with the formation of carbon dioxide gas.

Magnesium carbonate is used medicinally in stomach troubles.

MAGNESIUM CHLORIDE ($MgCl_2$)

Source in Nature.—Magnesium chloride is contained in sea water and in some mineral waters.

Preparation.—It may be prepared by the action of hydrochloric acid on magnesia, or on the carbonate.

Properties.—It is a very soluble and deliquescent salt. If present in water it corrodes metal boilers by the action of hydrochloric acid, which is liberated.

MAGNESIUM SULPHATE ($MgSO_47H_2O$). Epsom salts

Source in Nature.—This salt may occur in large deposits or in solution in certain mineral waters.

Preparation.—It can be prepared easily by the action of sulphuric acid on magnesia or magnesium carbonate.

$$MgO + H_2SO_4 = MgSO_4 + H_2O$$
$$MgCO_3 + H_2SO_4 = MgSO_4 + H_2O + CO_2$$

Properties.—Magnesium sulphate consists of small crystals with a slightly bitter taste and high solubility in water.

This salt is one of the best purgatives in the pharmacopœia, and is used extensively in bowel complaints.

Calcium (Ca). Atomic weight 40

Source in Nature.—The pure metal calcium is not found in nature, but in combination with carbon dioxide is abundant everywhere. The carbonate may be found in crystalline form as marble, and in the amorphous form as limestone and chalk.

Preparation.—It is derived from calcium chloride by electrolysis.

Properties.—It is a hard yellowish metal which, so long as the air is dry, does not tarnish on exposure, but if damp oxidizes at once with the liberation of hydrogen.

$$Ca + H_2O = CaO + H_2$$

The metal when heated burns very brilliantly.

Compounds

CALCIUM OXIDE (CaO). Quicklime

Preparation.—Commercially quicklime is prepared by roasting limestone, calcium carbonate, in lime kilms.

Properties.—It is a white amorphous substance which when damped with water swells and crumbles into a powder, giving off great heat. The powder resulting from the action is known as slaked-lime ($Ca(OH)_2$).

Quicklime is the basis of cement and mortar.

Slaked-lime is very slightly soluble in water, forming " lime water," which is alkaline in reaction. Whether as a solid or in solution lime freely combines with the carbon dioxide of the air to form chalk, calcium carbonate.

$$CaO + CO_2 = CaCO_3$$

CALCIUM CHLORIDE ($CaCl_2$)

Preparation.—Calcium chloride may be obtained by the action of hydrochloric acid on chalk or lime and evaporating to obtain crystals.

$$CaCO_3 + 2HCl = CaCl_2 + H_2O + CO_2$$

The crystals have 6 atoms of water of crystallization ($CaCl_26H_2O$), which can be removed by roasting, leaving an anhydrous (without water) and amorphous powder.

Properties.—The calcium chloride deprived of its water of crystallization is a white substance occurring in masses having a great affinity for water and so deliquescent in the presence of any dampness that it is used as a drying agent in laboratories.

CHLORIDE OF LIME ($Ca(OCl)Cl$). Bleaching Powder

Preparation.—Bleaching powder is prepared by exposing slaked-lime ($Ca(OH)_2$) to chlorine gas.

A considerable amount of chlorine is absorbed by the lime.

Properties.—Chloride of lime is a white alkaline powder with a strong smell of chlorine, which is present to the extent of 33 per cent. When stored in a damp atmosphere or added to water the powder decomposes liberating chlorine gas. It is this action, namely the freeing of chlorine, which makes it a very useful bleaching agent, and owing to its poisonous effect on organisms, it is used in surgical work as an antiseptic and even more so as a sterilizing chemical in the purification of water supplies. Dakin's solution, used in the treatment of gun-shot wounds, is a mixture of a solution of boric acid and bleaching powder.

CALCIUM CARBONATE ($CaCO_3$)

Source in Nature.—It is one of the commonest rocks forming the strata of the earth crust, and occurs in a crystalline state as marble or in an amorphous state as limestone or chalk.

Properties.—It is insoluble in water, but soluble when carbonic acid gas is present when it forms calcium bicarbonate.

It is the presence of the latter salt and a corresponding salt of magnesium, that is the cause of temporary hardness in water.

Acids act on calcium carbonate, liberating carbon dioxide gas.

$$CaCO_3 + 2HCl = CaCl_2 + H_2O + CO_2$$

If carbon dioxide gas is passed into lime water, for example, by breathing into it through a glass tube, the lime water will turn milky owing to the formation of a white precipitate of calcium carbonate.

Calcium carbonate as marble or lime stone is used for building purposes or by kilning in the manufacture of lime and cement.

As prepared chalk, it is used therapeutically as the basis of many powders.

CALCIUM CARBIDE (CaC_2)

Preparation.—Calcium carbide is made by heating quicklime (CaO) and coke in a furnace.

Properties.—This greyish-white substance, which occurs in hard masses, is used for the production of acetylene gas (C_2H_2), which is given off when water is added to the carbide.

$$CaC_2 + 2H_2O = Ca(OH)_2 + C_2H_2$$

CALCIUM SULPHATE ($CaSO_4$)

Source in Nature.—It occurs in rock form as alabaster and gypsum, both of which contain two molecules of water of crystallization ($CaSO_32H_2O$).

Properties.—When gypsum is roasted it loses some of the water molecules and becomes plaster of paris, and if the latter is ground into a powder and mixed into a paste with water it rapidly sets into a hard mass, gypsum being reformed.

Calcium sulphate is somewhat soluble in water, and it is this salt along with the magnesium sulphate which is the main cause of permanent hardness in water.

CALCIUM LACTATE

Calcium lactate, a compound used in pharmacy a great deal, is prepared by the action of lactic acid on calcium carbonate. It is a white almost tasteless powder.

CALCIUM PHOSPHATE

Occurs in nature, and is the chief constituent of bone ash. It is used as a manure.

CALCIUM SULPHIDE

When pure lime and sulphur are heated together, a sulphide of lime is produced, which after exposure to light becomes luminous, and so is used incorporated with paint to form luminous paint.

PRACTICAL CHEMISTRY—CHAPTER 17

EXPT. 98

Magnesium

Examine a piece of magnesium wire; it is light and has a silvery-white appearance.

Bend a piece of the wire; it does not fracture. Hammer it on a hard surface; the metal spreads without fracturing. It is therefore malleable.

Hold a piece in a pair of forceps and light one end of the wire; it burns with a dazzling white flame. Collect the white powder (magnesium oxide) formed.

EXPT. 99

Magnesium Carbonate

Compare the weight of heavy and light magnesium carbonate by lifting two bottles of similar size, one filled with the heavy, the other filled with light magnesium carbonate.

Shake up a pinch of the light magnesium carbonate in half a test-tube of water; it is insoluble. Add a little dilute sulphuric acid (H_2SO_4) to the tube; a gas escapes, and magnesium carbonate has exchanged its CO_2 radicle for SO_3 of the H_2SO_4, forming magnesium sulphate.

Do the same with the heavy magnesium carbonate. The same reaction occurs.

EXPT. 100

Magnesium Sulphate

Examine the crystals. Taste and note the bitter flavour.

Shake together equal parts of the crystals and water in a test-tube; it dissolves readily.

EXPT. 101

Calcium

Dip a knitting needle into a solution of calcium chloride and hold the wet end in a bunsen flame; the flame is coloured red—characteristic of calcium and its compounds.

EXPT. 102

Calcium Oxide

Quicklime as usually supplied is in lumps. Drop a few pieces of quicklime into a test-tube and a few drops of water. The tube gradually becomes warm, then hot; the lime swells and becomes powdery and may finally burst the tube.

Shake a pinch with a test-tube of water; it is slightly soluble. Test for reaction with litmus.

EXPT. 103

Calcium Chloride

Repeat Expt. 66.

Examine a piece of the white porous salt, taste; it has a warm, bitter taste.

Shake together in a test-tube one part of the salt and one part of water; it dissolves readily, the tube becoming hot.

EXPT. 104

Chloride of Lime

Examine the powder; it has a strong smell of chlorine and a gritty feel when rubbed between the fingers.

Shake some of the powder with water in a test-tube; it forms a milky solution which settles out, leaving a clear solution on top and a white insoluble powder of slaked lime below. Pour off the clear solution into a test-tube and dip into it a piece of red litmus paper; the paper first turns a blue colour and then white, showing that it is alkaline and possesses bleaching properties.

Add to the tube containing the clear solution a little dilute sulphuric acid, frothing occurs, and a heavy green gas (chlorine) collects in the tube. This shows that chlorine has dissolved from the bleach by merely shaking it with water.

EXPT. 105

Calcium Sulphate

Rub a little of the anhydrous powder (plaster of paris) into a paste with water; it rapidly sets to a hard mass due to its rehydration, when gypsum is again formed.

Add a pinch to half a test-tube of water and note that it is sparingly soluble.

Strontium (Sr). Atomic weight 87·6

Source in Nature.—Strontium is found as the carbonate in certain ores.

Preparation.—It is manufactured by the electrolysis of strontium chloride.

Properties.—It is a white metal much resembling calcium.

Compounds

By roasting strontium carbonate the oxide (SrO) is formed, and on addition of water to the latter the hydroxide is obtained.

STRONTIUM NITRATE ($Sr(NO_3)_2$)

Preparation.—It is made by dissolving the carbonate in nitric acid.

Properties.—It is used in the manufacture of fireworks to make red flares.

STRONTIUM BROMIDE

Is a salt sometimes used in medicine.

All strontium salts burn with a bright red colour.

Barium (Ba). Atomic weight 137

Source in Nature.—Barium occurs plentifully as barium carbonate and barium sulphate.

Preparation.—The metal barium is made from barium chloride by electrolytic action.

Properties.—Barium is a pale yellow, extremely heavy metal, and is not unlike calcium or strontium in its properties. It tarnishes when exposed to air, forming the oxide.

Compounds

BARIUM OXIDE (BaO). Baryta

Properties.—Barium oxide is produced by heating barium nitrate. It combines readily with water and carbon dioxide gas.

$$BaO + CO_2 = BaCO_3$$

On heating the oxide to a dull red heat it takes up another atom of oxygen and forms barium peroxide (BaO_2), but again loses the oxygen on heating still further. This is a method of obtaining oxygen on a large scale.

BARIUM HYDROXIDE ($Ba(OH)_2$)

Preparation.—Is formed by slaking the oxide with the addition of water.

Properties.—It is an alkaline solution which, like the calcium hydroxide, absorbs CO_2 to form barium carbonate.

$$Ba(OH)_2 + CO_3 = BaCO_3 + H_2O$$

BARIUM CARBONATE ($BaCO_3$)

Source in Nature.—It occurs as a native mineral.

Preparation.—In the pure form it is obtained by interaction of barium chloride and sodium carbonate.

$$BaCl_2 + Na_2CO_3 = BaCO_3 + 2NaCl$$

Properties.—It is a heavy white powder which, mixed with some bait, is used as a rat poison.

BARIUM CHLORIDE

Preparation.—Barium chloride is prepared by dissolving barium carbonate in dilute hydrochloric acid.

BARIUM NITRATE ($Ba(NO_3)_2$)

This salt is used in fireworks to produce a brilliant green colour.

BARIUM SULPHATE ($BaSO_4$)

Source in Nature.—Barium sulphate is the most important barium compound found in nature.

Preparation.—It can be prepared in the pure form by the action of sulphuric acid on any of the barium salts.

Properties.—It is insoluble in water and only very slightly so in weak acids. It is used as a white paint and for X-ray work.

Cadmium (Cd). Atomic weight 112

Cadmium metal is comparatively rare. It is used to make alloys and is employed in this combination in dentistry.

CADMIUM IODIDE

Is used along with starch for the detection of chlorine in water after treatment with " bleach." It is therefore one of the chemicals found in the Horrocks box.

CADMIUM SULPHIDE

Is an intensely yellow salt and is used to make yellow paint.

Zinc (Zn). Atomic weight 65

Source in Nature.—Zinc ores are found abundantly in various parts of the world. The two commonest compounds found are zinc sulphide and zinc carbonate.

Preparation.—The zinc ores are roasted, the sulphide to remove the sulphur and the carbonate to remove the carbon dioxide gas.

$$2ZnS + 3O_2 \text{ (heat)} = 2ZnO + 2SO_2$$
$$ZnCO_3 + \text{Heat} = ZnO + CO_2$$

Zinc oxide results in both cases and requires further reduction by roasting with coal dust.

$$2ZnO + C = 2Zn + CO_2$$

Properties.—Zinc is a hard somewhat brittle metal, bluish-white in colour, which tarnishes on exposure to a moist atmosphere, zinc oxide being formed.

Zinc alloyed with copper forms brass, and iron coated with zinc is known as galvanized iron. It is used extensively in this latter form for roofing, because once the oxide is formed on exposure to the air, acid vapours which may be present in the air have practically no effect on it.

Zinc is easily dissolved by weak acids with the liberation of hydrogen, in fact, this is a very common method of producing the gas.

$$Zn + 2HCl = ZnCl_2 + H_2$$

Zinc is also soluble in boiling caustic alkali solutions.

Compounds

ZINC OXIDE (ZnO). Chinese white

Preparation.—Zinc oxide is formed when zinc is strongly heated in air.

Properties.—It is almost insoluble in water.

It is used to make white paint; also mixed with lard forms zinc ointment, an antiseptic ointment much used as a wound dressing.

ZINC CARBONATE ($ZnCO_3$). Calamine

Source in Nature.—Zinc carbonate occurs as one of the ores.

Preparation.—A basic carbonate is formed by the interaction of a zinc salt with sodium carbonate.

Properties.—It is insoluble in water, though to some small extent if carbon dioxide gas is present; when roasted it gives off CO_2, being reduced to the oxide.

ZINC CHLORIDE ($ZnCl_2$)

Preparation.—It results from the action of hydrochloric acid and zinc.

$$Zn + 2HCl = ZnCl_2 + H_2$$

Properties.—Zinc chloride is a white very deliquescent substance, powerfully caustic in its action. It is used as a disinfectant, as a caustic for the removal of warts, and as a soldering flux.

ZINC SULPHATE ($ZnSO_4 7H_2O$)

Preparation.—It is produced when metallic zinc is dissolved in dilute sulphuric acid.

$$Zn + H_2SO_4 = ZnSO_4 + H_2$$

Properties.—It is a crystalline substance, the crystals resembling magnesium sulphate. It is very soluble in water. It is a strong antiseptic and is used extensively in surgery in the form of red lotion. It is also taken internally to stop diarrhœa, and liberated as a fine spray in a special chamber is used to disinfect the throats of soldiers in influenzal or cerebrospinal fever epidemics.

Radium

Radium is an extremely rare metal resembling barium, and possessing the property of giving off light in much the same manner as the X-rays. These " radiations," as they are called, are extremely powerful and can penetrate the tissues of the body. They have been found of great use in the treatment of certain diseases such as cancer.

Radium, or salts of radium, can impart this " radio activity " to other substances in close vicinity ; for instance, if a glass of distilled water is brought near a solution of a radium salt in a closed space the water will become radio active and give off rays of light, but loses the property rapidly if withdrawn from the closed space and exposed to air.

PRACTICAL CHEMISTRY—CHAPTER 17 (*continued*)

EXPT. 106

Zinc

Examine a piece of zinc. Note its bluish-white colour as compared with tin.

Hold a strip of the metal in the flame of a spirit lamp, it melts and burns with an intense blue flame forming a white powder—zinc oxide.

Repeat Expt. 59.

EXPT. 107

Zinc Oxide

Examine the powder. Rub between the fingers; it has a chalky feel.

Shake a little of the powder with water in a test-tube; it is almost insoluble.

EXPT. 108

Zinc Carbonate

Examine the powder (calamine); it is soft and nearly white.

Shake up some of the powder with a little water in a test-tube; it is insoluble.

EXPT. 109

Zinc Chloride

Handle with a pair of forceps and examine the salt; it usually turns a grey colour with age.

Shake up with some water in a test-tube; it is very soluble.

Expose a small piece to the atmosphere; it is deliquescent.

EXPT. 110

Zinc Sulphate

Mix equal parts of the crystals and water together in a test-tube, shake to dissolve. The crystals dissolve, showing that zinc sulphate is soluble in water.

QUESTIONS

QUESTION 1.—Name a soluble and insoluble salt of magnesium and calcium, and describe the properties of each.

QUESTION 2.—How are "quick" and "slaked" lime prepared? What happens when the former is exposed to air and the latter is mixed with water?

QUESTION 3.—Describe the preparation of chloride of lime. What action does it possess and to what uses is it put?

QUESTION 4.—From what source is zinc obtained, and what are its properties? Name two zinc compounds used for medicinal purposes and state on what special property their use depends.

CHAPTER 18

THE FIFTH GROUP—DIVALENT AND TRIVALENT ELEMENTS

Iron—Nickel—Cobalt—Chromium and Manganese

THE metals included in this group are very similar to one another, all being hard metals, some very much more so than others. They may be divalent or trivalent elements, thus forming two varieties of compounds, for example, ferrous and ferric, chromous and chromic, manganous and manganic, etc.

The " ous " salts are divalent—for example, ferrous oxide (FeO)—and the " ic " salts trivalent—for example, ferric oxide (Fe_2O_3).

Iron (Fe). Atomic weight 56

Source in Nature.—Iron in the pure state is found as small particles embedded in certain rocks and occasionally in large masses which have fallen to the earth as meteorites (solid or semi-molten rocks that fly about the sky and are known to the ordinary person as falling stars). In the combined state it may occur as many compounds, for example, as sulphides, carbonates, or oxides. It is from the latter two that the purest metal is obtained.

Preparation.—The oxide ore, chalk, and coal are heated in blast furnaces, and the molten iron is run out from time to time into sand moulds. It is converted by various processes into cast iron, wrought iron, and steel.

Properties.—The three kinds of iron mentioned above, namely, *cast iron, wrought iron,* and *steel,* are so totally unlike each other as regards their properties that they might easily be taken for different metals.

WROUGHT IRON

The purest form is " wrought iron," which is a greyish-white metal capable of taking a high polish. It is fibrous in texture and moderately hard when cold, but when hot can be moulded into any shape or form. Iron does not rust if exposed to dry air, but if the air is moist the oxide forms, which though forming a complete coating does not protect the surface of the iron from further corrosion.

It is soluble in hydrochloric and sulphuric acids, setting free hydrogen.

$$Fe + H_2SO_4 = FeSO_4 + H_2$$

In this form iron becomes fairly easily magnetized, but does not retain its magnetism.

STEEL

Steel is impure iron, varying amounts of carbon being present. Steel is harder than wrought iron, and can be made intensely hard by heating to redness and then plunging it into cold water. It can be softened by heating and then cooling slowly.

Steel is of crystalline nature and is a very strong metal, withstanding tremendous strain.

It is easily magnetized and retains its magnetism, and is used in the manufacture of magnets.

Stainless steel which stays bright is an alloy of steel and chromium but is softer than ordinary steel.

CAST IRON

This is the least pure of the three forms of iron, containing as it does large amounts of carbon and other impurities such as sulphur, phosphorus, and perhaps manganese.

Cast iron is brittle and can be easily broken. It is greyish-white in colour and in the molten condition can be moulded into any shape. One hardly need describe here the many and valuable uses to which these various forms of iron are put.

Compounds

Oxides.

These are three in number: ferrous, ferric, and magnetic oxide.

FERROUS HYDROXIDE ($Fe(OH_2)$)

Preparation.—It is prepared by the action of caustic soda on ferrous sulphate.

$$FeSO_4 + 2NaOH = Fe(OH)_2 + Na_2SO_4$$

Properties.—It readily takes up oxygen when exposed to air, changing its colour from white to red through shades of green to become the ferric oxide " rust."

Ferrous hydroxide dissolves in acids to form ferrous salts.

FERRIC OXIDE (Fe_2O_3). Rust

Origin in Nature.—It occurs plentifully in nature.

Properties.—It is a reddish-brown powder, which dissolves in acids to form ferric salts.

MAGNETIC IRON OXIDE (Fe_3O_4)

Origin in Nature.—Magnetic oxide of iron is found in nature as the mineral called magnetite.

Properties.—As found in the crude state it possesses magnetic properties which remain permanently in the ore. Magnetic oxide is dissolved by the action of sulphuric or hydrochloric acid into ferrous and ferric salts.

Ferrous Salts. The divalent salts of iron

The ferrous salts when heated in the presence of oxygen or treated with substances which give up their oxygen, readily unite with the oxygen to form ferric salts. Ferrous salts are therefore included in substances called " reducing agents," for they reduce compounds to a simpler form by the abstraction of oxygen.

FERROUS CARBONATE ($FeCO_3$). Green sand

Source in Nature.—Ferrous carbonate occurs abundantly in England, in fact, it is the commonest iron ore.

Properties.—It is soluble in water containing carbonic acid gas, and such waters may be found naturally in springs.

FERROUS SULPHATE ($FeSO_47H_2O$). Green vitriol

Preparation.—The only method of obtaining pure ferrous sulphate is by the action of dilute sulphuric acid on iron wire.

Properties.—It occurs in the form of light greenish-blue crystals which are soluble in water. When the seven molecules of water of crystallization are driven off by heating, the white amorphous anhydride, $FeSO_4$, which is very deliquescent, is left.

In both forms ferrous sulphate is used medicinally.

It is also present in small amounts combined with alum in the compound alumino ferric used in the clarification of water.

Ink is a mixture of ferrous sulphate and tannin.

FERROUS CHLORIDE ($FeCl_2$)

Is prepared by passing chlorine gas over iron filings.

It is not used medicinally.

FERROUS IODIDE

Is prepared by the action of iodine on iron filings. It is used medicinally as a syrup.

Ferric Salts. The trivalent salts of iron

FERRIC SULPHATE ($Fe_2(SO_4)_3$)

Preparation.—Ferric sulphate is prepared by the oxidation of the ferrous sulphate.

Properties.—It is a dark, brown, very deliquescent substance.

FERRIC SULPHIDE (FeS_2). Iron pyrites

Source in Nature.—This salt occurs in nature as " iron pyrites," and is mined rather for the sulphur than for the iron.

FERRIC CHLORIDE (Fe_2Cl_6)

Preparation.—Ferric chloride can be obtained by the action of hydrochloric acid on ferric oxide.

$$Fe_2O_3 + 6HCl = Fe_2Cl_6 + 3H_2O$$

Properties.—It is a dark-brown, very deliquescent substance, in solution forming a dark-brown liquid. It is used in medicine and is known as perchloride of iron.

Cyanides of Iron

Other iron salts include the ferrocyanides, the commonest in use being potassium ferrocyanide. It is a yellow crystalline substance, very stable, non-poisonous, and forming Prussian blue when ferric chloride is added to it.

It is used as a test agent for the presence of certain metals, giving a distinct colour reaction for each. Iron salts give a blue, copper a brown colour. Compound citrates of iron and ammonia and iron and quinine are used a great deal in pharmacy.

Nickel (Ni). Atomic weight 59

Source in Nature—Nickel occurs in nature as the arsenide and silicate and is associated with cobalt, copper, iron and manganese.

Preparation.—It is extracted by a complicated process when it is first roasted in a furnace and then freed from its impurities by treatment with various chemicals.

Properties.—Nickel is a silver-white hard metal which is not easily oxidized, and because of these qualities it is used extensively in the electroplating of machinery, cutlery, etc.

With copper and zinc it forms an alloy, which is intensely white and is known as German silver. Apart from these industrial uses it is of no particular importance.

Cobalt (C). Atomic weight 59

Source in Nature.—It occurs along with nickel as the sulphide and arsenide, and is associated with the same metallic impurities as nickel.

Preparation.—It is extracted with difficulty by the application of great heat and subsequent chemical action.

Properties.—In the pure state cobalt much resembles nickel and iron, but is harder than the latter.

Its salts are distinguished by their prominent colours, the blue silicate of cobalt being used in glass staining, etc.

Manganese (Mn). Atomic weight 55

Source in Nature.—Manganese occurs in nature as manganese dioxide (MnO_2).

Preparation.—Extracted by treating the ores in a blast furnace.

Properties.—It is very similar to cast iron, being brittle. It rapidly oxidizes, and therefore must be kept in sealed tubes or under naphtha. In the metallic state it is of no practical value.

Compounds of Manganese

OXIDES OF MANGANESE

There are four oxides—Manganese (MnO), Manganic (Mn_2O_3)—and two others, which act like the higher oxide of chromium in forming manganates, permanganates.

MANGANESE DIOXIDE. Pyrolusite

Source in Nature.—This compound is very abundant in nature.

Properties.—When heated, manganese dioxide gives off some of its oxygen, and is therefore ranked as an oxidizing agent. Its action on hydrochloric acid results in the evolution of chlorine. This is the usual method of preparing chlorine gas.

Other manganese salts are the sulphate and carbonate.

POTASSIUM PERMANGANATE ($KMnO_4$)

Preparation.—By either adding water to a solution of potassium manganate or, as is more commonly done, by passing chlorine gas or carbon dioxide through the solution potassium permanganate is formed.

Properties.—It occurs as small, dark, purplish crystals, which are soluble in water, forming a violet-coloured solution.

It is one of the recognized oxidizing agents, giving up its oxygen readily on heating. Potassium permanganate is extensively used as a disinfectant.

PRACTICAL CHEMISTRY—CHAPTER 18

Epxt. 111

Iron

Examine a piece of iron; it is heavy and hard. Hold one end in a flame; heat travels quickly along the metal, which becomes hot to the hand, showing it is a conductor of heat.

Sprinkle a few iron filings on to a flame and note that they sparkle.

Expt. 112

Wrought Iron

Drop a piece of wrought iron (horseshoe) on to a stone floor; it does not fracture.

Repeat Expt. 52. It does not retain magnetism but is capable of being attracted.

Expt. 113

Steel

Examine a piece of steel (table knife blade). Although it is thinned down to a fine blade it can be bent without breaking, and springs back on being released.

Examine a piece of watch spring; it is hard, springy, and has a high polish with a blue colour.

Heat one end of the spring in a flame to a dull, red heat; remove and allow to cool slowly. If bent the spring will not now resume its former position, showing that the steel is now soft.

Again heat the softened end of the spring to a dull red and plunge it into some cold water; it has now been retempered and is springy again.

Expt. 114

Cast Iron

Drop a piece on to a hard floor, it will break up into a number of pieces, showing that cast iron is brittle. Examine one of the fractured ends and note the fine granular, grey, irregular surface of the fractured part.

Expt. 115

Ferric Oxide

Leave a few clean iron nails in water or hydrogen peroxide overnight; take out and dry and scrape off the hydrated ferric oxide or rust formed.

Shake up the powder with a little water; it is insoluble. Add a little dilute acid to the mixture; the rust dissolves, showing that the hydrated ferric oxide is soluble in dilute acids.

Expt. 116

Ferrous Sulphate

Examine the crystals, noting light green colour and irregular shape.

Expt. 117

Potassium Permanganate

Examine and note the dark purplish long needle-shaped crystals.

Shake some of the crystals with water and note that they are moderately soluble, colouring the water a deep purple colour. The solution does not change the colour of red or blue litmus paper.

Gently heat a few crystals in a test-tube; the crystals break up, liberating oxygen. Plunge a lighted match into the mouth of the tube; the flame burns with increased brilliance, indicating the presence of oxygen.

QUESTIONS

Question 1.—Name the main points of difference between wrought iron, cast iron and steel.

Question 2.—Iron is a divalent and trivalent element. What do you mean by this ?

Give an example of a divalent and trivalent iron compound, stating its properties and uses.

Question 3.—Name two compounds of manganese, stating their source or preparation and their properties.

CHAPTER 19

THE SIXTH GROUP—TRIVALENT ELEMENTS

Aluminium and Boron

THE substances in this group being trivalent, require three atoms of any monovalent element before they will unite to form compounds. Aluminium will therefore require three times the number of atoms of the monovalent element, chlorine, before the compound aluminium chloride ($AlCl_3$) can be formed.

Aluminium (Al). Atomic weight 27

Source in Nature.—The pure metal is not found in nature but is plentiful in combination with silicon as clay. It may also be present in volcanic rocks.

Properties.—Aluminium is a white metal, comparatively soft, and is one of the lightest of the metals. Aluminium does not tarnish; even with the addition of heat very little oxidation takes place.

Cold hydrochloric acid and sulphuric acid dissolve it, the former more readily than the latter. Dilute nitric acid when cold acts slowly, but when boiling has a fairly vigorous action on the metal. Caustic soda also dissolves it readily.

Compounds of Aluminium

ALUMINIUM HYDRATE ($Al_2(OH_3)$)

This substance when dissolved in water forms a gelatinous precipitate which slowly settles and is used as a means of clarifying water supplies. It is also used as a fixing agent in dye works. Dyed cloth so treated can be washed and the colour does not wash out.

ALUMINIUM SULPHATE ($Al_2(SO_4)_3 18H_2O$)

Preparation.—Aluminium sulphate may be extracted from clays containing this metal, or can be obtained in the pure form by the action of sulphuric acid on aluminium hydrate.

Properties.—It dissolves freely in water, forming an acid solution. This salt has the same uses as the hydrate, and is the chief constituent in the Army clarifying powder. In the manufacture of aluminium sulphate there may be present as an impurity a small amount of ferrous sulphate. When this occurs, the compound is known commercially as alumino ferric.

The Alums

These are *double* salts of aluminium and another metal. The two most important are *potash alum* and *ammonia alum.*

POTASH ALUM. Potassium aluminium sulphate, or alum

Properties.—It is a colourless crystalline substance acid in reaction. It dissolves slightly in cold water and readily in hot water. In fact, boiling water will dissolve about one hundred times more alum than water at freezing point. Burnt alum is obtained by heating potash alum and driving off the water of crystallization. It acts as a " fixing agent " for dyes, and is used because of its astringency as a drug in medicine.

ALUMINIUM SILICATE. Clay

Source in Nature.—Aluminium silicate occurs everywhere as clay, the purity of the salts depending on the particular clay. The coarser varieties of clay are used for the manufacture of bricks and earthenware utensils, while with the purer forms, by adding substances to assist in the binding together of the clay particles, various qualities of pottery are made. The process in both cases is carried out by baking the articles in a furnace.

Boron (B). Atomic weight 11

Source in Nature.—Boron does not occur in the pure state, but as borax and boric acid.

Preparation.—Boron is extracted from the oxide by heating with potassium or sodium.

$$B_2O_3 + 6Na = 3Na_2O + 2B$$

The substance obtained in this way occurs as an amorphous powder and requires to be heated to a white heat with aluminium before the crystalline form of boron can be obtained.

Crystalline boron is a very hard substance.

The test to demonstrate the hardness of substances is to scratch the one with the other, the harder will always scratch the softer. The diamond will scratch any other substance and is used in glass cutting. Boron crystals will scratch any other substance except a diamond.

Compounds

BORON TRIOXIDE (Ba_2O_3)

Combines with water to form boric acid (H_3BO_3).

BORIC ACID

Source in Nature.—Boric acid occurs in nature dissolved in the waters which issue as streams from fissures in the ground in certain volcanic districts.

Boric acid crystals are purified by dissolving them and allowing them to crystallize out afresh.

Properties.—The white shining flake-like crystals of boric acid have a soft and soapy feel.

They are much more soluble in hot water than in cold water.

A solution of boric acid has an acidity equal to that of carbonic acid. It is not sufficiently acid to turn litmus completely red, but it alters it to a port wine colour. The presence of boric acid can be demonstrated by its action on turmeric paper, which is stained red on drying.

Boric acid burns with a green colour when heated in the bunsen flame.

Boric acid is a valuable antiseptic.

BORAX

Is a sodium salt of boric acid which is found in nature and can be prepared by the addition of sodium carbonate to a boiling solution of boric acid. The large colourless crystals which result on cooling the mixture are very slightly alkaline in reaction.

It is used as a soldering flux, as a weak alkali, and in the manufacture of vita glass, a special glass that allows the passage of ultra-violet rays.

PRACTICAL CHEMISTRY—CHAPTER 19

Trivalent

EXPT. 118

Aluminium

Examine a piece of the metal (cooking vessel). Note the colour and weight.

Hold one end in the flame; heat travels along the metal slowly, showing that aluminium is a moderate conductor of heat. Keep the metal in the flame; it does not melt.

Mix some filings with HCl or NaOH in a test-tube; the metal dissolves with evolution of hydrogen.

EXPT. 119

Alum

Shake together one part of water with three parts of alum in a test-tube; it is sparingly soluble.

Heat the tube and the solubility will be seen to increase rapidly. Continue heating to boiling point when the whole of the alum will dissolve. Dip a piece of blue litmus paper into the solution; the paper turns red, showing that alum dissolved in water is acid.

Now heat a few crystals on a tin lid. They first melt in their water of crystallization. Continue heating; the moisture boils off, and the salt is now a white porous mass known as burnt alum.

EXPT. 120

Boric Acid

Rub a little of the powder between the fingers; note the soapy feel.

Add a pinch of the powder to half a test-tube of water and shake; it is almost insoluble. Now heat the tube and contents; the powder slowly dissolves. Dip a piece of turmeric paper in the solution, allow the paper to dry and add a drop of a solution of caustic soda to the red stain. It turns black. This is a test for boron.

Sprinkle a little of the powder on to the flame of a bunsen burner or spirit lamp. Note the green colour.

QUESTIONS

QUESTION 1.—What are the natural sources and properties of aluminium ?

Describe the action that takes place when the sulphate of aluminium is added to water and state one of the uses of this salt.

QUESTION 2.—State what you know of the element boron, and its compound boric acid.

CHAPTER 20

THE SEVENTH GROUP—DIVALENT AND TETRAVALENT ELEMENTS

Lead and Tin

Lead (Pb). Atomic weight 206

Source in Nature.—Lead is most frequently found as the sulphide galena.

Preparation.—Metallic lead is extracted from the natural sulphide by smelting.

Properties.—It is a soft heavy metal capable of being scored by the finger nail. When cut with a knife lead has a shining bluish-white surface. It tarnishes rapidly, however, on exposure to damp air, forming on its surface a film of the oxide which protects the lead from further corrosion.

When heated to a high temperature oxide of lead is formed. Nitric acid alone has any vigorous action on lead, producing lead nitrate.

Some organic acids will dissolve small amounts of lead, and this is of real importance if it occurs in lead pipes supplying drinking water. If the water is soft the hydrate of lead is formed, which is insoluble unless some peaty acid is present. In this way the consumers may be slowly poisoned by absorbing soluble lead into their systems. Should chalk be present as in hard waters, the insoluble lead carbonate will be formed which will coat the inside of the pipe and prevent further erosion. Because of its softness and malleability, the ease with which it melts and its resistance to most corrosive agents, it is used for many purposes, for example, roofing, pipes, rifle bullets, etc. Alloyed with tin it forms solder and pewter, which, before the era of china, was the metal used in the manufacture of plates, dishes and tankards.

Compounds of Lead

The Lead Oxides.—There are several oxides, the best known being lead oxide (PbO) and red lead (Pb_3O_4).

LEAD OXIDE (PbO). Litharge

Preparation.—Lead oxide is formed when lead is strongly heated in the presence of air or oxygen.

Properties.—It is a heavy yellow powder which fuses on heating and when allowed to cool forms a crystalline mass called litharge.

It is very slightly soluble in water and may be made into a plaster or an ointment for external use in skin affections.

RED LEAD (Pb_3O_4)

Preparation.—Red lead is obtained by heating the carbonate of lead to a temperature of about 400° C.

Properties.—It occurs as a scarlet powder which is slightly soluble, but which is readily decomposed by dilute acids forming lower oxides and lead salts. Red lead is used as a basis for red paint, a common preliminary coating to iron work.

LEAD CHLORIDE ($PbCl_2$)

Preparation.—It results from the action of hydrochloric acid on lead.

Properties.—It is a white crystalline substance highly soluble in hot, but almost insoluble in cold water.

LEAD SULPHIDE (PbS)

As has been shown, occurs in the native form galena.

LEAD CARBONATE ($PbCO_3$)

Preparation.—Occurs in nature associated with galena, and is found coating the interior of lead pipes, when the water supply is hard.

Properties.—When powdered finely is known as white lead and is the basis of most white paints.

LEAD IODIDE (PbI_2)

Preparation.—Lead iodide is prepared by the interaction of lead acetate and potassium iodide.

Properties.—It consists of a heavy bright yellow powder or scaly crystals almost insoluble in cold, but soluble in hot water. It is the basis of certain lead ointments used externally for skin diseases.

LEAD ACETATE. Sugar of lead

Preparation.—Lead acetate is prepared by dissolving lead oxide or lead carbonate in acetic acid.

Properties.—It exists as small white crystals, has a sweet taste, and is poisonous. It is the lead salt contained in most of the lead pills and lotions used in medicine.

Tin (Sn). Atomic weight 118

Source in Nature.—Tin is not found in the metallic state, but generally as the oxide (SnO_2) in certain mineral ores.

Properties.—Tin is a shining white metal which does not tarnish, remaining bright when exposed to air. It can be cut with a knife, being harder than lead but softer than zinc. It is malleable and can be beaten into thin sheets known as tinfoil.

The metal when strongly heated is converted into the oxide (SnO_2).

Tin is employed in the process of " tinning," which consists of coating other metals with a thin film by dipping them into molten tin.

Tin is commonly used to make alloys ; with lead it combines to form pewter and solder, with copper it occurs as brass and bronze.

PRACTICAL CHEMISTRY—CHAPTER 20

Expt. 121

Lead

Examine a piece of lead. Note its weight. Bend it with the fingers and hammer it on a hard surface. It is soft and malleable. Scrape the surface with a knife and note the shining bluish metal beneath, which will quickly become dull.

Hold a piece in the flame ; it readily melts.

Expt. 122

Tin

Heat a piece of so-called tin sheeting (iron dipped in tin), and while hot rub the heated part with a cloth. The tin coating is removed, exposing the iron sheeting beneath.

Leave a piece of the tin plate in water overnight ; it does not rust.

QUESTIONS

Question 1.—Give briefly the main characteristics of the metals tin and lead.

CHAPTER 21

EIGHTH GROUP—TETRAVALENT ELEMENTS

Carbon and Silicon

Carbon (C). Atomic weight 12

Mention has been made in a previous chapter of the allotropic modification of carbon, namely, the diamond, graphite, and charcoal.

They are identical substances, yet it is very difficult to believe this considering the difference in their appearance, the diamond being a brilliantly clear crystal, graphite a scaly black shining substance, and charcoal a black soft amorphous powder. Carbon also occurs in combination with many elements ; in the form of carbon dioxide gas, it is a constant constituent of the air, and with hydrogen it occurs as marsh gas, while the number of carbonates found in nature are legion, the most common being chalk or limestone, the carbonate of calcium.

DIAMOND

Source in Nature.—It is found in various parts of the world, especially in South Africa and South America. Being somewhat scarce the diamond is numbered as one of the most precious stones.

How the diamond is formed is still unknown and all attempts to manufacture it satisfactorily by laboratory methods have failed.

It can be converted by heat into its allotropic forms, graphite and carbon, but the reverse action is still to be discovered.

The proof that the diamond consists of pure carbon has been shown by burning the diamond in oxygen by means of an electric lamp ; no residue is left, the diamond is changed to black carbon and then combines with oxygen to form carbon dioxide, the presence of which can be demonstrated by allowing this gas to mix with lime water when a milky solution results.

The diamond does not conduct electricity ; this is not an important point in itself, but as graphite does readily and charcoal only slowly, it shows that, although the same substance, these allotropic forms have very distinct and definite properties.

Diamonds are mined because of their rarity and are sold as gems. They are also used for cutting glass.

GRAPHITE (Black lead)

Source in Nature.—It is found far more abundantly than the diamond, and is widely distributed in all parts of the world.

Preparation.—It can be made by roasting coke with molten iron in the requisite amounts.

Properties.—Graphite is a soft blackish substance occurring as crystalline flakes, which have a smooth or soapy feel.

When strongly heated it burns, forming carbon dioxide gas.

It is used for the lead in pencils and also as blacking with which grates are rubbed in order to protect the iron from corrosion.

It is also used on account of its greasiness as a lubricant in place of oil. Unlike the diamond it is a good conductor of electricity, and is used to make electrodes.

AMORPHOUS CARBON. Charcoal

Preparation.—Carbon in varying degrees of purity can be prepared by the destructive distillation of organic substances, for example, wood, bone, coal, etc.

When a log of wood burns in the fire, nothing is left of it but a white ash, all the carbon has passed up the chimney as carbon dioxide gas, or adheres to the inside of the chimney as soot. If the amount of carbon lost as carbon dioxide and soot is reduced to a minimum by limiting the air getting to the wood when it burns, then instead of wood ash there remains a black brittle substance which in shape and size resembles the logs from which it was derived.

This substance is charcoal and the method of its preparation is called "the destructive distillation of wood." The charcoal burner obtains charcoal by building a pile of wood faggots, and surrounding them with peat or earth sods. He sets fire to the wood through a vent and the whole mass slowly smoulders without being reduced to an ash, any gases which are formed passing up through an opening at the top. When all combustion has ceased the pile is opened up and the faggots are found to be converted into charcoal.

Animal charcoal can be obtained in the same manner by the destructive distillation of bones, the process usually taking place in large ovens or retorts. Coke is the carbon residue left in the manufacture of coal gas.

Charcoal derived from these sources is not pure carbon; other substances are present. An almost pure carbon can, however, be obtained by burning crude oil, turpentine, or any other substance rich in carbon with a limited amount of air present. Soot or lamp black results from the combustion and can be collected in large amounts, and is used in the preparation of black paints. In laboratories or in hospitals where tracings of the pulse and heart-beats are made, a special paper blackened with lamp black is required. The process of blacking the paper is carried out by first covering the vent at the base of a bunsen burner with some wool soaked in turpentine, and turning on and lighting the gas. Turpentine will be drawn up into the bunsen flame, which will smoke heavily, and the special paper when exposed to the flame will become rapidly blackened.

Properties.—The chief property possessed by charcoal is its power of absorbing gases. Any gas from oxygen to the most offensive emanation is absorbed. It is therefore called a deodorant, because it will remove the bad odours given off by decomposing or putrifying matter.

The container of the modern gas mask is filled with carbon, which acts as a filter and holds back any gas that may be encountered in war, and so protects the lungs and eyes. Not only will charcoal remove gases from the air, but it will also remove substances from solution.

For this reason it may be used as a filter for drinking water, all the objectionable matter being held back on passing through the charcoal. It may also be used as a decolorizing agent by means of which certain colouring matters can be removed from solution.

Charcoal is a bad conductor of electricity. One more important property is possessed by carbon, that is, its reducing power on certain metal compounds. At high temperature carbon breaks down the compounds of sodium, potassium, and phosphorus, with the freeing of these metals in their pure state. In fact, it is almost universally used in the smelting process of metallic ores, acting not only as a fuel for the furnace, but as a reducing agent.

Compounds of Carbon

These may be divided into organic and inorganic compounds. The organic comprise those compounds which go to build up animal and vegetable matter. They are combinations of carbon, hydrogen, oxygen, and sometimes of nitrogen, and their number is endless. What is considered necessary for the student to know will be found in the organic section of this book.

The division between organic and inorganic substances is very fine—in fact, it is probable that the one class of carbon compounds merges into the other. Inorganic compounds of carbon may have acquired their carbon originally from an organic source. Animals and vegetables give off carbon dioxide when alive, and when dead and decomposing break up into simple compounds of which carbon dioxide is one. This carbon dioxide mixes with the air or becomes dissolved in water containing inorganic substances with which it unites to form carbonates.

For example, lime water is a solution of an inorganic substance, and if one blows into it, it becomes milky due to carbon dioxide of the breath uniting with the lime in the water to form calcium carbonate.

Calcium carbonate is considered an inorganic compound, but as shown in this case, part of it may have an organic origin.

It is generally understood, however, that inorganic salts of carbon are those formed apart from living matter, either vegetable or animal.

A rough and ready way of classifying carbon compounds into their two main groups, is to take all those containing hydrogen as organic and those without hydrogen as inorganic.

Oxides of Carbon

There are two oxides, carbon monoxide and carbon dioxide (carbonic acid gas).

CARBON DIOXIDE (CO_2). Carbonic acid gas

Source in Nature.—Carbon dioxide occurs in small quantities in the pure state as a constant constituent of air, and is abundant in combination with many elements. Limestone and chalk are the commonest natural compounds of carbon dioxide.

Preparation.—It can be easily obtained from calcium carbonate, either by heating as was described under quicklime, or by the action of such an acid as hydrochloric. The two methods can be shown by the formulae

$$CaCO_3 + \text{Heat} = CaO + CO_2$$

and

$$CaCO_3 + 2HCl = CaCl_2 + H_2O + CO_2$$

The gas can be collected in the latter method by allowing it to bubble into an inverted glass cylinder filled with water which the gas replaces.

Properties.—It is a colourless gas with a faint sweetish odour, similar to that which one smells on opening a bottle of soda water, which is really a solution under pressure of carbon dioxide.

When carbon dioxide is subjected to a great pressure or its temperature is reduced to about 80° below freezing point, it becomes a clear colourless liquid. When this liquid is allowed to escape from a cylinder, some of it will at once return to the gaseous state, but in doing so, it gives up the cold that it absorbed when liquefying, and this cold is so intense that it freezes the rest of the liquid as it emerges.

The frozen carbon dioxide is known as carbon dioxide snow or "dry ice," and it can be cut with a knife to form a pencil with which certain skin diseases can be treated. The gas is soluble in water and is the cause of temporary hardness, because the insoluble carbonates become soluble when there is sufficient carbon dioxide present to form the bicarbonates.

The gas can be dissolved in a much greater amount than is normally present by increasing the pressure. This fact is made use of in the manufacture of aerated waters.

The gas is given off as a waste product of respiration and any atmosphere will become polluted if the room which man or animals occupy is not properly ventilated. Fires, lamps, gas, etc., all add to the total of carbon dioxide present, so that the amount present in the air of a room is a guide to the purity or otherwise of the air in the room.

There is as little as 0·03 per cent. present in pure air. When it increases to 3 per cent., which it may do in certain trades, for example, brewing, then symptoms are felt, and if it is as concentrated as 6 per cent. then it kills the person by suffocation.

Carbon dioxide does not burn but remains as a product of combustion.

Certain metals, such as sodium, potassium and manganese, when roasted in carbon dioxide abstract the oxygen leaving pure carbon.

$$4Na + CO_2 = Na_2O + C$$

Its non-inflammable nature has made it a recognized method of extinguishing fires. The gas is blown on to the fire, which it at once puts out, as nothing can burn in its presence.

CARBONIC ACID (H_2CO_3)

Preparation.—Carbonic acid is a solution of carbon dioxide in water.

$$CO_2 + H_2O = H_2CO_3$$

Properties.—Carbonic acid is a weak acid, resembling boric acid in this respect; it changes litmus to a plum colour instead of to a bright red as do the stronger acids.

It is very unstable, that is, it soon breaks up into water and carbon dioxide gas. This occurs when carbonic acid is either boiled or frozen.

$$H_2CO_3 + \text{Heat} = H_2O + CO_2$$

Carbonic acid forms salts called carbonates.

Carbonates

These are formed by the action of carbonic acid on various metals.

Two kinds of salts are formed, the carbonates and the bicarbonates.

Take, for example, the compounds of calcium and carbonic acid, there is the ordinary carbonate or chalk ($CaCO_3$), which is insoluble, but which, when excess of carbonic acid is present in water, becomes the soluble acid carbonate or bicarbonate, the cause of temporary hardness in water. The reason why it is called temporary hardness is because it can be removed by boiling, which drives off the excess carbon dioxide, leaving as a sediment the insoluble carbonate.

$$\underset{\substack{\text{Soluble}\\\text{bicarbonate}}}{CaCO_3H_2CO_3} - CO_2 = \underset{\text{Insoluble carbonate}}{CaCO_3 + H_2O}$$

The carbonates of the alkaline metals may be formed by the direct action of carbon dioxide gas on the hydroxide.

$$2NaOH + CO_2 = Na_2CO_3 + H_2O$$

Other elements are made to combine with carbon dioxide to form carbonates by the interaction of a carbonate and a salt of the element.

$$CuCl_2 + Na_2CO_3 = CuCO_3 + 2NaCl$$

All carbonates can be decomposed on heating with the liberation of CO_2.

$$Na_2CO_3 = Na_2O + CO_2$$

The important carbonates are found mentioned with their respective elements.

CARBON MONOXIDE (CO)

Preparation.—Carbon monoxide does not exist normally in nature; it results from the combustion of carbon where the air or oxygen in which it is burnt is insufficient to oxidize fully the carbon to carbon dioxide, thus :

$$C + O_2 = CO_2$$

would become with a scarcity of oxygen

$$2C + O_2 = 2CO$$

This may occur in dug-outs heated with braziers or in garages where cars are left running, the exhaust gases containing carbon monoxide, or in machine-gun emplacements, after two or three thousand rounds of ammunition have been fired. It may also occur in mines where a fire breaks out, and there is not sufficient air to form carbon dioxide.

Carbon monoxide and hydrogen are produced in large quantities when steam is passed through red-hot charcoal.

$$H_2O + C = CO + H_2$$

It can be prepared by heating certain organic compounds, such as oxalic acid with strong sulphuric acid, but is mixed with certain amounts of carbon dioxide.

The carbon dioxide can be removed by bubbling the mixture of gases through lime water.

Properties.—Carbon monoxide is a colourless gas and is extremely poisonous, causing severe symptoms even in very small amounts. It kills by suffocation at a concentration of 0·3 per cent. It is extremely dangerous as its presence cannot be detected by the sense of smell. It burns with a blue flame, which is often visible at the top of a coal fire, and carbon dioxide is produced.

$$2CO + O_2 = 2CO_2$$

Carbon monoxide is feebly soluble in water and passes through lime water without turning the latter milky. Like carbon it is a powerful reducing agent, absorbing oxygen from metallic oxides and freeing the pure metal. This is one of its main uses in industry. It is present to the extent of about 7 per cent. in ordinary coal gas.

The Hydrocarbons

The hydrocarbons are compounds of hydrogen and carbon, and although they are all essentially organic compounds, nevertheless, it is customary to deal with three of them, namely, marsh gas (CH_4), and ethylene (C_2H_4) and acetylene (C_2H_2) in the inorganic section.

MARSH GAS (CH_4)

Source in Nature.—This gas occurs generally where putrefaction, especially of vegetable matter, is taking place. In stagnant foul-smelling ponds, bubbles of the marsh gas may be seen bursting on the surface. It has been the cause of many disasters in coal mines, the miners calling it " black damp." Coal was once masses of vegetable matter, trees, peat, etc., which through pressure and the passing of millions of years, have been converted into coal. In the early stages of its formation putrefactive gases were evolved, and they may have become locked up in cavities in the coal seam and be tapped by the miners excavating.

Preparation.—It may be prepared in the laboratory by heating a mixture of sodium acetate and caustic soda.

Properties.—Marsh gas is a light colourless gas which burns, but does not give light when doing so.

It explodes when burnt along with air or oxygen, and the resulting carbon dioxide is known to the miners as " after-damp."

ACETYLENE (C_2H_2)

This gas may be formed when marsh gas and coal gas are burned with a supply of air insufficient to get complete combustion. Its presence can be detected by its characteristic smell. This is sometimes observed when a bunsen burner is lit by mistake at the bottom of the chimney. It is prepared readily by the action of water on calcium carbide.

$$CaC_2 + 2H_2O = Ca(OH)_2 + C_2H_2$$

Properties.—Acetylene gas has an offensive and penetrating odour, and when lit, burns with a brilliant white light.

Its main use is for lighting purposes.

ETHYLENE (C_2H_4)

This gas occurs as one of the hydrocarbons in coal gas.

Preparation.—It can be prepared by the action of concentrated sulphuric acid on alcohol.

$$C_2H_5OH = C_2H_4 + H_2O$$

The gas is collected in jars inverted over water.

Properties.—It burns with a bright luminous but smoky flame and explodes when burned with oxygen or air.

Silicon (Si). Atomic weight 28

Source in Nature.—The element silicon is not found in nature in the pure state, but in combination is one of the principal constituents of most rocks.

Compounds of Silicon

SILICON DIOXIDE (SiO_2). Silica

Source in Nature.—Silica in the comparatively pure form of quartz and flint or compounded with other substances, forming granite, clay and many rocks, is found widely distributed in all parts of the world.

Silicates

There are many salts of silicon, of which the alkaline silicates are soluble in water and go by the name of water-glass.

CLAY

Clay is composed of aluminium silicate, and various grades of it are used in the manufacture of bricks, earthenware and china. Purer forms, such as fullers' earth and kaolin, are employed in medicine, the one as a dusting powder and the other in the form of a poultice.

GLASS

Glass is manufactured from compounds of silicon. There are many varieties of glass, the difference depending on the nature of the salts that enter into its composition.

The commoner are potassium, sodium and calcium, and according to the proportion of these present so the glass will be hard or soft.

The necessary amounts of sand, potassium or sodium carbonate, lime and lead are heated in a furnace until a fused mass results.

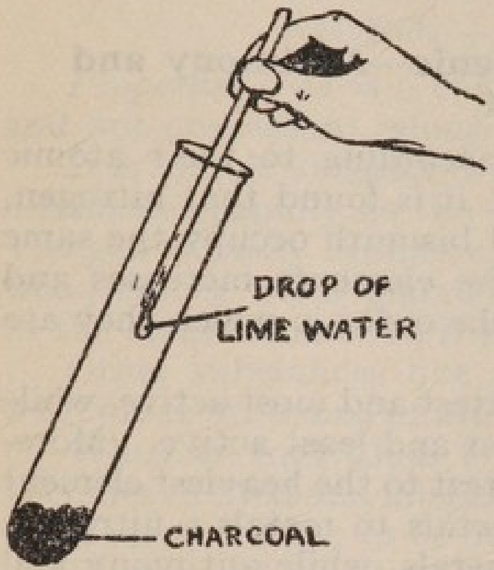
Hard glass withstands great heat before fusing, whereas soft glass fuses readily with moderate heat. The harder glass is used in the manufacture of laboratory apparatus which has to withstand heat. Soft glass, when in a fused condition, can be blown and is therefore used in making bottles, flasks, etc. The only substance which has any pronounced action on glass is hydrofluoric acid, which therefore must not be kept in a glass container. This acid is used in marking the graduations on glass burettes, etc.

Glass may be coloured by the addition of certain metallic oxides : those of cobalt producing blue ; chromium, green ; and copper, red.

PRACTICAL CHEMISTRY—CHAPTER 21 (*continued*)

Expt. 123

Carbon



(i) Show examples of carbon, vegetable and animal charcoal, coke and graphite.

(ii) Remove the lead (graphite) from a pencil and heat a portion in a test-tube until all gum has been driven off. Remove and powder the lead placing it in another test-tube.

Heat the powder to red heat, and introduce into the test-tube a glass pipette which has been dipped into fresh lime water. The lime water becomes opaque due to the presence of CO_2.

Expt. 124

Carbon Dioxide

(i) Heat some powdered marble or chalk in a test-tube, and test for CO_2 as in the previous experiment.

(ii) To some marble or chalk add some dilute hydrochloric acid, and note the bubbles of CO_2 coming off. Test as above.

QUESTIONS

Question 1.—Describe the three allotropic forms of carbon. What proof is there that they are identical substances ?

Question 2.—How is vegetable charcoal prepared, and what properties and uses has it ?

Question 3.—What are the sources and properties of carbon dioxide and carbon monoxide ?

Question 4.—State what you know of marsh gas. To what group of substances does it belong ?

CHAPTER 22

THE NINTH GROUP—TRIVALENT AND PENTAVALENT ELEMENTS

Nitrogen—Phosphorus—Arsenic—Antimony and Bismuth

In the classification of elements according to their atomic weights, as described in Chapter 12, it is found that nitrogen, phosphorus, arsenic, antimony, and bismuth occupy the same group. The atomic weight of these elements increases and their chemical action decreases in the order in which they are named.

Nitrogen compounds are the lightest and most active, while bismuth compounds are the heaviest and least active. Moreover, as one progresses from the lightest to the heaviest element the substances change from non-metals to metals ; nitrogen, phosphorus, and arsenic are non-metals, while antimony and bismuth are definite metals. There is a corresponding change in the reaction of the salts formed from the oxides, those of nitrogen being strongly acid, while those of bismuth are basic or alkaline.

Nitrogen (N). Atomic weight 14

Source in Nature.—Nitrogen is the chief constituent of air, of which it comprises four-fifths.

It occurs combined as nitrates, and is present in many organic compounds.

Preparation.—On dissolving in an alkaline pyrogallate solution, the oxygen and carbon dioxide from the air, that which remains is practically pure nitrogen.

Properties.—It is a most inactive substance, in fact there is nothing at all distinctive about it. It is a colourless, tasteless, odourless, and neutral gas, which remains unaffected, unless exposed to intense heat, when it can be made to unite with oxygen to form the reddish-brown gas nitrogen peroxide.

Certain organisms living on the ground have the power of combining nitrogen with hydrogen to form ammonia gas. Nitrogen is only feebly soluble in water, is non-poisonous, and will not support combustion.

Compounds of Nitrogen

There are five oxides of nitrogen, of which two are of importance, namely nitrous oxide (N_2O) and nitrogen pentoxide

(N_2O_5), partly because of their own qualities and partly because of the qualities of the acids formed from them.

NITROUS OXIDE (N_2O). Laughing gas

Preparation.—Nitrous oxide is usually prepared by heating ammonium nitrate.

$$NH_4NO_3 + \text{Heat} = 2H_2O + N_2O$$

Properties.—It is a colourless gas, having a sweetish taste and not unpleasant odour.

It is soluble in water to a moderate extent, but its solubility diminishes rapidly as the temperature of the water is raised. The gas is easily broken up, giving off oxygen, and a glowing wood ember when placed in it will at once burst into flame in the same manner as if placed in oxygen.

Other substances like phosphorus and sulphur will burn with great brilliance in nitrous oxide.

When an individual breathes the gas for a little time he becomes excited and laughs heartily. This effect on the person has caused the gas to be named "laughing gas."

If the administration of the gas is continued, unconsciousness and later death will result. Unconsciousness occurs rapidly and easily, and for this reason nitrous oxide is a very suitable form of anæsthetic, especially in dentistry. When used as an anæsthetic it has to be thoroughly purified after being prepared as described. Nitrous oxide can be liquefied comparatively easily.

NITRIC ACID (HNO_3). Aqua fortis

This acid corresponds to nitrogen pentoxide (N_2O_5), which is a white crystalline substance very unstable and very soluble in water, forming nitric acid.

Preparation.—Nitric acid is prepared in large quantities from either sodium or potassium nitrate (saltpetre) by the action of sulphuric acid.

$$2NaNO_3 + H_2SO_4 = Na_2SO_4 + 2HNO_3$$

Properties.—Nitric acid is an oily colourless liquid, which when exposed to the air gives off irritating fumes. It is a very powerful acid, corroding most metals, and it absorbs moisture readily from the air. Nitric acid is a powerful oxidizing agent, setting free its oxygen readily, and because of this it rarely gives off its hydrogen as other acids do when acted on by a metal, but oxidizes the hydrogen to form water.

Nitric acid does not act on platinum or gold, but when it is mixed with hydrochloric acid, forming what is called aqua regia, then these metals are dissolved. Combined with metals nitric acid forms nitrates.

NITROUS ACID (NHO_2)

This substance does not occur in the pure state, but is mixed with varying quantities of nitric acid. It decomposes rapidly into nitric acid, nitric oxide, and water.

$$3HNO_2 = HNO_3 + 2NO + H_2O$$

It acts as a reducing agent, taking up oxygen to form nitric acid.

Nitrous acid forms nitrites when combined with metals.

It is an opportune moment to describe here the chemistry of the air, of which nitrogen is the chief constituent.

The Atmosphere or Air. Density 14·5

The air, which is a mixture of gases, surrounds the world and is absolutely necessary for the existence of animal and vegetable life.

The gases that enter into its composition, and the comparative amounts that are present, are as follows :

(1) Oxygen	20·93
(2) Nitrogen.	79·04
(3) Carbon dioxide	0·03

In addition there are varying amounts of moisture present, and minute quantities of the "rare gases." Of the mixture oxygen is necessary for life, the nitrogen has no function except that of a diluent, that is, it dilutes the strength of oxygen which in the pure state would be too stimulating ; and the carbon dioxide is absorbed by plants, the carbon abstracted and oxygen set free into the air. Air obeys all the laws of gases as described in Chapter 5. It is not visible, and its presence can only be appreciated when it is set in motion, being felt in varying degrees from a slight breeze to a powerful hurricane.

Phosphorus (P). Atomic weight 31

Source in Nature.—Phosphorus does not occur pure in nature, but is found combined in certain metals which when they crumble and form earth, give it to plants, and the plants hand it on to the animals that feed on them, so that nearly all plants and all animals contain phosphorus.

In animals the chief source is the bones, where the phosphorus is found combined with calcium to form calcium phosphate.

Preparation.—Phosphorus is prepared from bone ash, first by the action of sulphuric acid, when a soluble acid phosphate is formed, and then after the solution has been evaporated to a thick syrup, by heating it with charcoal, when the phosphorus distils over.

Properties.—It is a pale yellow transparent substance, in form not unlike bees-wax.

When the temperature is reduced to freezing point it becomes brittle. On raising the temperature slightly above normal it is easily melted, and when heated strongly it boils, giving off vapour. When phosphorus is exposed to air it will be found to give off light in the dark, showing that slow combustion is occurring. It is a very inflammable substance, bursting into flame at a very low temperature, or when it is cut with a knife, the friction being sufficient to set it on fire. In fact, so liable is it to go on fire, that any cutting should be done under water, in which it should always be stored to prevent the slow combustion which occurs on exposure to air.

Phosphorus is not soluble in water, but is freely soluble in carbon disulphide. Phosphorus combines actively with oxygen and the halogens.

When a mixture of phosphorus and potassium chlorate or nitrate is pounded it explodes with violence. This fact is made use of in the manufacture of matches.

Phosphorus may exist as a red amorphous powder known as *red* or *amorphous phosphorus*. This is an allotropic modification of phosphorus, from which it is prepared by prolonged exposure to heat. Its properties are entirely different from those of yellow phosphorus. It does not oxidize in air, and therefore does not require to be immersed in water, nor is it luminous. It requires considerable heat to melt it or cause it to burn, and whereas yellow phosphorus is poisonous the red is harmless to man. The main use of phosphorus is in the manufacture of matches. The modern match is dipped in a thick liquid containing potassium chlorate, which is allowed to dry. The striking-paper on the box is coated with a mixture containing red amorphous phosphorus. The matches will not light unless struck on the specially prepared paper, and therefore they are called safety matches.

PHOSPHORIC ACID (H_3PO_4)

Preparation.—It is prepared by dissolving phosphorus pentoxide in water, or by the action of sulphuric acid on burnt bones.

Properties.—It is moderately acid in reaction, forming phosphates when combined with metallic elements.

Phosphates, especially calcium phosphate, occur in bone ash and are used as manure for fields. The iron phosphate is included in the tonics called Easton's Syrup and Parrish's Food.

Hypophosphites

These are salts also used in medicine.

Arsenic (As). Atomic weight 75

Source in Nature.—Arsenic is occasionally found in the free state, but generally combined with iron and sulphur.

Preparation.—It can be prepared by heating arsenious oxide, As_4O_6, with charcoal.

$$As_4O_6 + 6C = 6CO + 4As$$

On a large scale it is obtained by roasting the mineral ore in a special furnace provided with chambers, in which the vapours of arsenic can condense.

Properties.—Arsenic is a non-metal, and may exist in three allotropic forms, two of which are crystalline and the third amorphous.

In the crystalline form it is a shiny greyish brittle substance, which readily sublimes when the temperature is raised, and which burns with a bluish flame with the formation of arsenious oxide. Arsenic is a very poisonous substance.

Compounds

ARSENIURETTED HYDROGEN (AsH_3)

Similar to phosphorus and nitrogen, arsenic unites with hydrogen.

Preparation.—Arseniuretted hydrogen can be prepared by introducing arsenious oxide into a mixture, usually zinc and hydrochloric acid, from which hydrogen is being generated.

Properties.—Arseniuretted hydrogen is a colourless evil-smelling gas, and is one of the most poisonous substances known. It burns with a bright blue flame, forming arsenious oxide and water.

When strongly heated, for example, by a bunsen burner as the gas passes through a glass tube, amorphous arsenic is deposited as a mirror on the cooler parts of the tube. This is what is termed Marsh's test for the presence of arsenic.

ARSENIOUS OXIDE (As_4O_6)

This is the substance popularly known as arsenic.

Preparation.—When arsenic pyrites is roasted with carbon, arsenious oxide is deposited.

Properties.—It is a coarse white crystalline substance, somewhat resembling china in appearance. It is very slightly soluble in water, and is exceedingly poisonous.

It is from this oxide that most arsenical compounds are formed, and it used to enter into the composition of many paints, especially those of green colour.

It is used medicinally, and is the arsenical compound present in Fowler's solution.

ARSENIOUS ACID (H_3AsO_3)

Preparation.—It is prepared by dissolving arsenious oxide in water. It forms salts called arsenites, of which sodium arsenite is used as a fly poison and copper arsenite, Scheeles green or Paris green, as a poison to kill mosquito larvæ in water.

The other oxide is *arsenic oxide* from which *arsenic acid* is formed by dissolving it in water.

This acid forms *arsenates*, of which *sodium arsenate* is the most commonly used in medicine. There are two complex arsenical compounds which must be mentioned although no attempt is made to describe their chemical composition. These are *salvarsan* 606 and *neo-salvarsan* 914, both most important drugs in the treatment of syphilis.

Antimony (Sb). Atomic weight 122

Source in Nature.—It occurs combined with sulphur in antimony pyrites.

Properties.—Antimony is a very brittle light crystalline substance which burns brilliantly when heated. It is alloyed with lead and tin to form " type metal " for the printing press.

There are three oxides, but only one need be considered, namely, *antimonious oxide* (Sb_4O_6), which is used as a drug in medical treatment, as also is its combination with potassium tartrate, which goes by the name of tartar emetic.

Bismuth (Bi). Atomic weight 209

Source in Nature.—The metal may occur in the pure form, but is generally found as the oxide or sulphide in mineral ores.

Preparation.—The ores containing the metal are roasted, when bismuth can be separated in pure form.

Properties.—It is a shiny white metal which does not tarnish when exposed to dry air, but does if the air is moist.

It forms alloys readily with other metals, and confers the qualities of hardness and easy fusibility to them.

Compounds

It combines with the halogens to form bismuth chloride, iodide, and bromide.

There are four oxides, and if any one of these is acted on by an acid, the salt of that acid is formed. The salt of bismuth used in medicine is the carbonate, which substance is opaque to the X-rays, and therefore can be taken in the form of a " meal " to demonstrate the position of the various parts of the alimentary canal.

PRACTICAL CHEMISTRY—CHAPTER 22

Expt. 125

Composition of the Air

Mark a test-tube in five equal divisions. Into a cup place a dessert-spoonful of pyrogallic acid from the photographic dark-room and also about a table-spoonful of a 10 per cent. solution of caustic soda.

Turn the test-tube upside down with its mouth under the brown liquid and leave it. After some time the liquid will rise to the first mark, showing that one-fifth of the air in the tube is oxygen. The remainder is practically all nitrogen, but being an inert gas there is no means of demonstrating its presence.

Expt. 126

Properties of Nitrogen

Light a match, turn the mouth of the above test-tube upwards, insert the match. It goes out. All the oxygen has been absorbed by the pyrogallic solution and the nitrogen left will not support combustion.

EXPT. 127

Preparation of Nitric Acid

Mix a little saltpetre and sulphuric acid in a test-tube; fumes of nitric acid are given off.

QUESTIONS

QUESTION 1.—How are nitrous oxide and nitric acid prepared? Describe the properties of each.

QUESTION 2.—What is the composition of the air, and of what importance are the main constituents?

QUESTION 3.—What is the source of phosphorus and how prepared? Describe the properties of its two allotropic forms.

QUESTION 4.—Arsenic is a non-metal. What do you understand by this? How is it prepared, and how is its presence indicated?

QUESTION 5.—How is arsenious oxide prepared? Describe its properties and actions.

CHAPTER 23

TENTH GROUP—DIVALENT NON-METALLIC ELEMENTS

Oxygen and Sulphur

THE members of this group unite with hydrogen, oxygen to form water (H_2O), and sulphur to form sulphuretted hydrogen (H_2S).

Oxygen (O). Atomic weight 16

Oxygen is the most important element that exists, for without it all life would cease.

Source in Nature.—Oxygen occurs abundantly, both in the pure state as the most important constituent of air, and in combination with other elements, for example, in water and in many of the metallic ores.

Preparation.—Oxygen is most readily prepared for laboratory purposes by heating potassium chlorate (mention of which has been made under potassium), when the whole of the combined oxygen is set free and potassiun chloride remains.

$$2KClO_3 = 2KCl + 3O_2$$

On a large scale oxygen is manufactured by liquefying air and allowing it to evaporate from the nitrogen.

Properties.—Oxygen is a colourless, tasteless, and odourless gas, which is slightly soluble in water, but on raising the temperature to boiling point all oxygen is driven off. Its presence in water is just as essential for the fishes that live there as its presence in the air is to the animals that inhabit the earth.

Combustible substances like wood, coal, and candles will not burn unless sufficient oxygen is present, and such substances burn more brightly and vigorously the greater the proportion of oxygen present.

In pure oxygen, glowing carbon bursts into flame, and phosphorus and sulphur burn brightly with the formation of carbon dioxide, phosphorus pentoxide, and sulphur dioxide gases.

$$C + O_2 = CO_2$$
$$P_4 + 5O_2 = 2P_2O_5$$
$$S + O_2 = SO_2$$

Oxygen gas is stored in metal cylinders, and is used in cases of partial suffocation resulting either from inflammation of the lung or of some asphyxiant such as coal gas having been inhaled.

Compounds of Oxygen

The commonest oxide is the oxide of hydrogen—namely, water, which has been described under hydrogen. Practically every element except fluorine forms oxides ; some are acid in reaction, turning blue litmus red, others are alkaline or basic, turning red litmus blue.

Examples of *basic oxides* are :

Sodium potassium and calcium oxides.
Copper oxide.
Barium oxide.
Manganese oxide.
Iron oxide.
Lead oxide.

Of *acid oxides :*

Sulphur dioxide.
Phosphorus pentoxide.
Carbon dioxide.
Nitrogen pentoxide.

Of *neutral oxides :*

Certain oxides of iron, manganese, and lead.

Ozone (O_3). An allotropic form of oxygen

Source in Nature.—Ozone is present in the air near the sea coast, and is probably due to the effects of electricity present in the air on oxygen.

Preparation.—It is best prepared by the action of a silent electric discharge on oxygen. If the discharge is allowed to spark, then the ozone is destroyed as fast as it is formed.

Properties.—It is a gas heavier than air, with a faint, peculiar smell not unlike that of chlorine.

It is an oxidizing agent, the extra molecule of oxygen being given off when ozone comes in contact with certain metals.

$$2Ag + O_3 = 2Ag_2O + O_2$$

Ozone slowly breaks up into oxygen and is therefore an oxidizing agent, bleaching materials and killing germs.

It is used on the Continent as a means of purifying drinking water.

Sulphur (S). Atomic weight 32

Source in Nature.—Sulphur occurs free in the vicinity of volcanoes, and in some parts of the world actually exists in deposits of considerable thickness.

The metals lead, iron, copper, zinc, mercury, and some others are found combined with sulphur in the form of sulphides. Thus galena is the sulphide of lead, iron pyrites and copper pyrites the sulphides of iron and copper.

Preparation.—As crude sulphur is usually mixed with mineral impurities, these have to be got rid of, and this is done by melting the sulphur, allowing it to drain off.

From the metallic ores, sulphur is obtained by roasting; thus copper or iron pyrites in the process of smelting yield considerable amounts of sulphur dioxide.

$$2FeS_2 + 11O = Fe_2O_3 + 4SO_2$$

Sulphur can be got from sulphur dioxide by the action of hydrogen sulphide (H_2S) on it,

$$SO_2 + 2H_2S = 2H_2O + 3S$$

or the sulphur dioxide can be used for the manufacture of sulphuric acid (H_2SO_4).

Whichever method is employed the crude sulphur must be purified, and this is done by heating it until it boils, when sulphur vapour is given off, and this is condensed in cold chambers into a soft powder consisting of minute crystals and going by the name of " flowers of sulphur." If this powder is heated it melts, forming a brownish-coloured liquid, which is run into moulds and allowed to set into blocks or rods. Sulphur in this form is known as " roll sulphur."

Properties.—Sulphur is a pale yellow, brittle, crystalline solid, insoluble in water but freely soluble in carbon disulphide and to a lesser degree in chloroform and turpentine.

When heated sulphur first melts and becomes a clear brown liquid, then as more heat is applied it becomes darker and thicker, and finally with increasing heat it becomes thin again and boils, giving off sulphur vapour.

As it cools it goes through the same stages, but in the reverse manner. Sulphur has four allotropic modifications, two of them differing in the shape of their crystals, the third being of the nature of rubber, and the fourth an amorphous powder.

Sulphur is used in the treatment of many skin diseases, especially scabies.

HYDROGEN SULPHIDE (H_2S)

Source in Nature.—Hydrogen sulphide is found in the region of volcanoes; it is also formed on the decomposition of vegetable and animal matters and is present dissolved in the water of sulphur springs.

Preparation.—It is usually prepared by the action of sulphuric or hydrochloric acid on ferrous sulphide (FeS).

$$FeS + H_2SO_4 = FeSO_4 + H_2S$$

Properties.—Hydrogen sulphide is a colourless gas, having an extremely offensive odour similar to rotten eggs.

It is very poisonous when present in large amounts in the air. There have been many fatal accidents to men working in sewers due to their entering a section of the sewer where this gas has accumulated.

It is moderately soluble in water, and burns with a blue flame with the formation of sulphur dioxide and water.

$$2H_2S + 3O = 2SO_2 + 2H_2O$$

Hydrogen sulphide tarnishes metals such as silver, tin, and lead, and owing to its behaviour with certain metallic salts in solution, it is used in laboratories for the identification of the metal present.

SULPHUR DIOXIDE (SO_2)

Source in Nature.—Sulphur dioxide is present in the air near volcanoes, in certain mines, and also in the air of towns, from the smoke of coal fires. During a London green fog the irritability to the throat is due to this gas, and its presence can be identified by the tarnishing of silver and brass articles in the house.

Preparation.—Sulphur dioxide gas is obtained when sulphur burns in air.

$$S + O_2 = SO_2$$

It also results from the roasting of iron and copper pyrites, as has been shown under sulphur.

The best laboratory method consists of heating copper with sulphuric acid, when the gas is evolved, copper sulphate and water being formed.

$$Cu + 2H_2SO_4 = CuSO_4 + 2H_2O + SO_2$$

Properties.—Sulphur dioxide gas is a colourless gas with an intensely pungent smell. It does not burn nor will anything burn in it; a burning match will immediately go out. It is a heavy gas, nearly twice as heavy as air, and therefore it should be let loose near the ceiling when it is used in the disinfection of rooms.

It dissolves readily in water, forming sulphurous acid.

$$SO_2 + H_2O = H_2SO_3$$

Sulphur dioxide gas can be liquefied by reducing the temperature and increasing the pressure, and it is employed in this form contained in a syphon as a handy method of liberating sulphur dioxide in the disinfection of a room, or for the destruction of bugs, rats, and other pests. It is an active bleaching agent, destroying certain vegetable colours, and is used in this rôle quite considerably and especially with articles that would be destroyed by the action of chlorine. It is also used to remove excess chlorine left in articles bleached by chloride of lime.

Sulphurous Acid (H_2SO_3)

Is a solution of sulphur dioxide gas in water, and is an acid and bleaching agent.

Sulphuric Acid (H_2SO_4). Oil of vitriol

Preparation.—Sulphuric acid is manufactured in immense quantities by the oxidation of sulphur dioxide gas in the presence of water. The chemical used to supply the sulphur dioxide with oxygen is nitric acid. During the action, nitric acid is converted first into nitrogen peroxide, then to nitric oxide, which gas has a great affinity for oxygen and takes it up from the air, so that there is an unlimited supply of oxygen.

The following formula shows the main stages of the chemical action.

$$(1)\quad SO_2 + \underset{\text{Nitric acid}}{2HNO_3} = H_2SO_4 + \underset{\text{Nitrogen peroxide}}{2NO_2}$$

$$(2)\quad SO_2 + \underset{\text{Nitrogen peroxide}}{NO_2} + H_2O = H_2SO_4 + \underset{\text{Nitric oxide}}{NO}$$

$$(3)\quad \underset{\text{Nitric oxide}}{2NO} + \underset{\text{from air}}{O_2} = \underset{\text{Nitrogen peroxide}}{2NO_2}$$

Properties.—Sulphuric acid is a thick, oily, colourless fluid without smell. When it is added to water the mixture gets very hot and may splash violently out of the vessel which contains it. Care therefore must be taken when mixing the two to add the sulphuric acid drop by drop to the water.

Sulphuric acid absorbs moisture from the air or from any damp substance; for this reason it is employed in laboratories as a drying agent. It even abstracts water from its combination with certain organic substances, leaving a charred mass, which is the carbon of the compound. For example, if sulphuric acid is allowed to act on sugar, an organic substance, the water is abstracted and the sugar blackens, showing the presence of carbon.

$$C_6H_{12}O_6 + H_2SO_4 = C_6 + (H_2SO_4 + 6H_2O)$$

Sulphuric acid is employed in many chemical factories, and is also used to produce nitric and hydrochloric acids. Diluted, it is used as an astringent in certain stomach troubles.

Sulphuric acid forms sulphates when it acts on certain elements, and these will be found described under these substances.

SODIUM THIOSULPHATE, OR HYPO

Is a salt which is used in photography, and also sometimes to remove the taste of chlorine from a water that has been sterilized by bleaching powder.

CARBON DISULPHIDE (CS_2)

Preparation.—This substance is prepared by passing sulphur vapour over red-hot charcoal.

$$C + S_2 = CS_2$$

Properties.—Carbon disulphide is a colourless thin liquid with a distinctive odour.

It burns with a blue flame.

It is a solvent for various substances such as phosphorus, sulphur, fats, and, most important of all, rubber.

PRACTICAL CHEMISTRY—CHAPTER 23

EXPT. 128

Oxygen

Heat some potassium chlorate powder in a test-tube until it boils, then light a splinter of wood from a pencil and blow out the flame, leaving a red glow. Introduce the glowing splinter into the test-tube and note that it immediately bursts into flame.

EXPT. 129

Sulphur

Heat some flowers of sulphur in a test-tube; the sulphur darkens and finally melts, giving off fumes of sulphur which condense on the upper and cooler part of the tube.

EXPT. 130

Sulphur Dioxide

(i) Burn some sulphur in a tin lid; the pungent gas, SO_2, comes off.

(ii) Into a test-tube containing a little strong sulphuric acid drop some pieces of copper wire. Heat to boiling and note action on copper commences with the production of SO_2 gas.

EXPT. 131

Sulphuric Acid

Add some sulphuric acid drop by drop to a test-tube half filled with water and note the heat generated.

Add to a pinch of sugar some strong sulphuric acid and lay aside. The sugar gradually blackens, due to the abstraction of water, the carbon being left.

QUESTIONS

Question 1.—State the natural sources, mode of preparation, and properties of oxygen.

Question 2.—What is the usual source of sulphur ? In what forms does it appear and how is it prepared ? Briefly state its chief properties.

Question 3.—Compare the two gases—sulphur dioxide and sulphuretted hydrogen.

Question 4.—How is sulphuric acid prepared ? Describe its appearance and briefly state its action.

CHAPTER 24

THE ELEVENTH GROUP—MONOVALENT NON-METALLIC ELEMENTS. THE HALOGENS

Chlorine—Bromine—Iodine—Fluorine

THE members of the Halogen Group in many respects are very similar to one another. For instance, they all combine with hydrogen to form their respective acids—hydrochloric acid from chlorine and hydrobromic acid from bromine—and these acids are unlike any other acids in having no oxygen present in their combination. Sulphuric acid (H_2SO_4) and nitric acid (HNO_3) possess considerable amounts of oxygen, as will be seen from their formulae, but the acids of the halogens have none, for instance, HCl, HBr, HI.

The halogens differ in their weight and according to their weight, so they vary in the particular form they assume and in their chemical activity.

Chlorine is the lightest, is a gas and is the most active, turning bromine and iodine out of combination with other substances. Thus, when chlorine is brought into contact with starch and potassium iodide mixture, the chlorine unites with the potassium setting free iodine, which reveals its presence by turning the starch a blue colour.

$$KI + Cl = KCl + I$$

This is the chemical action which takes place in the Horrock's test for the presence of chlorine in water to which bleaching powder has been added.

Bromine is the next lightest of the halogens ; it is a liquid, and is less active than chlorine but more so than iodine. It will therefore only remove iodine from its combination with other substances.

The salts formed by the union of any one of the halogens and the same element are very similar in appearance.

Chlorine (Cl). Atomic weight 35·45

Source in Nature.—This gas does not occur free in nature, but is abundant combined with sodium, forming rock salt.

Preparation.—There are several methods of preparing chlorine. The one most commonly employed, when the gas is manufactured on a large scale, is by the action of manganese dioxide on hydrochloric acid in the presence of heat.

$$MnO_2 + 4HCl = 2H_2O + MnCl_2 + Cl_2$$

Another method is by the electrolysis of brine, and this is a method of generating chlorine for the purification of water supplies.

Properties.—Chlorine is a yellow-green gas, having a very pungent and unpleasant odour, and if inhaled is most irritating to the air passages. Like most other gases, it can be combined into either a liquid or solid form by altering the temperature and pressure. Chlorine gas does not burn, and when it is mixed with hydrogen and exposed to sunlight it combines with the latter, causing an explosion when doing so. The affinity chlorine has for hydrogen is very great, and if mixed with water, in which it is very soluble, it will slowly break up the water into hydrogen, with which it combines, and oxygen, which is set free and is seen to form bubbles in the fluid.

$$2H_2O + 2Cl_2 = 4HCl + O_2$$

Chlorine gas has a corroding effect on most metals, and therefore the cylinders in which it is stored should preferably be of stoneware. If the gas is dry, however, it can be kept in steel cylinders without risk of corrosion ; stored in this manner it is employed in the purification of water, the required dose being regulated by a special machine. The gas may be stored in lime, forming chloride of lime, and this is used commercially as a bleaching agent and also as a very cheap and efficient chemical in the purification of water. The bleaching action can only occur in the presence of water. If cloth or calico is soaked in a solution of chloride of lime, the chlorine set free breaks up the water, as has been described, with the liberation of oxygen, and it is the oxygen that removes the colour from certain dyes by the process of oxidation.

HYDROCHLORIC ACID

Source in Nature.—Hydrochloric acid occurs normally in small amounts in the air near certain volcanoes.

Preparation.—It can be prepared by the direct union of chorine and hydrogen, as already described, or by the commoner method of treating common salt (sodium chloride) with sulphuric acid ; sodium sulphate is formed and hydrochloric acid is set free.

$$2NaCl + H_2SO_4 = Na_2SO_4 + 2HCl$$

It is this action which takes place in the commercial manufacture of sodium sulphate where large quantities of hydrochloric acid solution are obtained as a waste product.

Properties.—Hydrochloric acid is a colourless gas, which forms thick white fumes when exposed to moisture. It has a pungent smell and is very acid in reaction, and dissolves in water, forming an acid solution.

Watery solution of hydrochloric acid

When a solution of hydrochloric acid is exposed to air it fumes, that is hydrochloric acid gas is set free, and this action is increased by heating the fluid. The acid dissolves certain metals, especially zinc and iron, with the freeing of hydrogen and the formation of the chlorides of these metals.

$$Zn + 2HCl = ZnCl_2 + H_2$$

Bromine (Br). Atomic weight 80

Source in Nature.—Bromine is never found free in nature, but as a compound with sodium, potassium, or magnesium. It occurs in sea water and certain mineral springs.

Preparation.—It is prepared in the same way as chlorine, by the interaction of manganese dioxide and potassium bromide in the presence of sulphuric acid and heat. Manganese and potassium sulphates and water are formed as a result of the chemical action, and the vapours of bromine condense as a deep red liquid.

$$2KBr + 2H_2SO_4 + MnO_2 = MnSO_4 + K_2SO_4 + 2H_2O + Br_2$$

It has already been noted that chlorine will displace bromine from its combination with a substance.

$$2KBr + Cl_2 = 2KCl + Br_2$$

Properties.—It is a dark-red heavy liquid, and the only non-metallic element that is a liquid at normal temperatures. It evaporates very rapidly, and a bottle which contains it will be seen to be filled with red bromine vapour above the level of the liquid. Bromine has a powerful and unpleasant smell, and is very irritating to the eyes and air passages. Bromine is used in the preparation of bromides.

HYDROBROMIC ACID (HBr)

Preparation.—This acid cannot be so easily prepared as hydrochloric acid by the direct union of bromine and hydrogen, but requires the addition of a considerable amount of heat.

The usual mode of preparation is by the action of phosphorus on bromine in the presence of water.

Properties.—Hydrobromic acid is colourless and has a pungent odour. It fumes when exposed to air and is highly soluble in water, forming an acid solution with properties very similar to hydrochloric acid.

Salts of Bromine

POTASSIUM BROMIDE (KBr)

This compound occurs as colourless crystals, which are very soluble and have a saline taste. It is used extensively as a drug in the treatment of nervous disabilities.

Other salts used in medical treatment are sodium, ammonium and strontium bromide.

Iodine (I). Atomic weight 127

Source in Nature.—It agrees with chlorine and bromine in not occurring free in nature. It is found combined with sodium, potassium, magnesium, and calcium in certain sea-weeds, and as an impurity in Chili saltpetre (sodium nitrate).

Preparation.—It is prepared from sea-weed by burning the weed to an ash, known as kelp, then extracting the soluble salts.

After further treatment, which frees certain impurities and converts the salts of iodine into hydriodic acid, the liquid is dosed with manganese dioxide and sulphuric acid, when iodine is formed and is condensed in containers.

$$2HI + MnO_2 + H_2SO_4 = MnSO_4 + 2H_2O + I_2$$

Properties.—Iodine is a blue-black solid, occurring as crystalline flakes not unlike graphite. At a temperature just above boiling point of water, iodine vaporizes, giving off a purple vapour.

The smell of iodine is not so pungent and irritating as chlorine.

It is only slightly soluble in water, but is freely soluble in potassium iodide solution, in alcohol, and ether. Iodine, even when present in the very smallest amounts, will turn a solution of starch a deep blue colour.

HYDRIODIC ACID (HI)

Preparation.—It requires considerable heat before hydrogen and iodine can be made to unite to form hydriodic acid. It is usually obtained by the action of phosphorus on iodine in the presence of water.

Properties.—Hydriodic acid is a colourless gas which fumes on exposure to air and is very soluble in water.

Salts of Iodine

POTASSIUM IODIDE

Is the principal salt that is used in medicine. It consists of white opaque crystals having a saline taste and being soluble in water. Mention of this salt has been made under potassium; similarly other iodides are mentioned under their respective metals.

Fluorine (F). Atomic weight 19

This element is so active that it is very difficult to preserve it in the free and gaseous state. It even attacks glass and is used in the form of hydrofluoric acid to mark graduations on burettes, flasks, etc.

PRACTICAL CHEMISTRY—CHAPTER 24

Expt. 132

Chlorine

Mix in a test-tube a pinch of pure sodium chloride and manganese dioxide or potassium permanganate. Pour over the mixture some dilute sulphuric acid and warm.

Chlorine comes off as a yellowish-green gas, the pungent smell of which is typical.

(Smell very cautiously.)

Expt. 133

Bromine

Repeat last experiment, using potassium bromide instead of sodium chloride. Brown fumes of bromine appear above the liquid, and may be present in sufficient quantities to condense (on the inside of the tube) to a dark brown liquid.

(Smell bromine fumes cautiously.)

Expt. 134

Iodine

Repeat experiment, this time using potassium iodide. Iodine vapour comes off and sublimes on the upper part of the tube.

Scrape off a little of the sublimate and dissolve in water in a test-tube; add some starch solution and note blue colour as evidence of iodine, or dip a pipette in starch solution, and note blueing of starch drop when introduced inside the test-tube.

QUESTIONS

Question 1.—What substances belong to the halogen group of elements, and in what respects do they differ from one another ?

Question 2.—Describe the source, preparation, and properties of chlorine.

Question 3.—Compare iodine with bromine. What are the two salts of these substances most commonly used in medicine ?

PART II

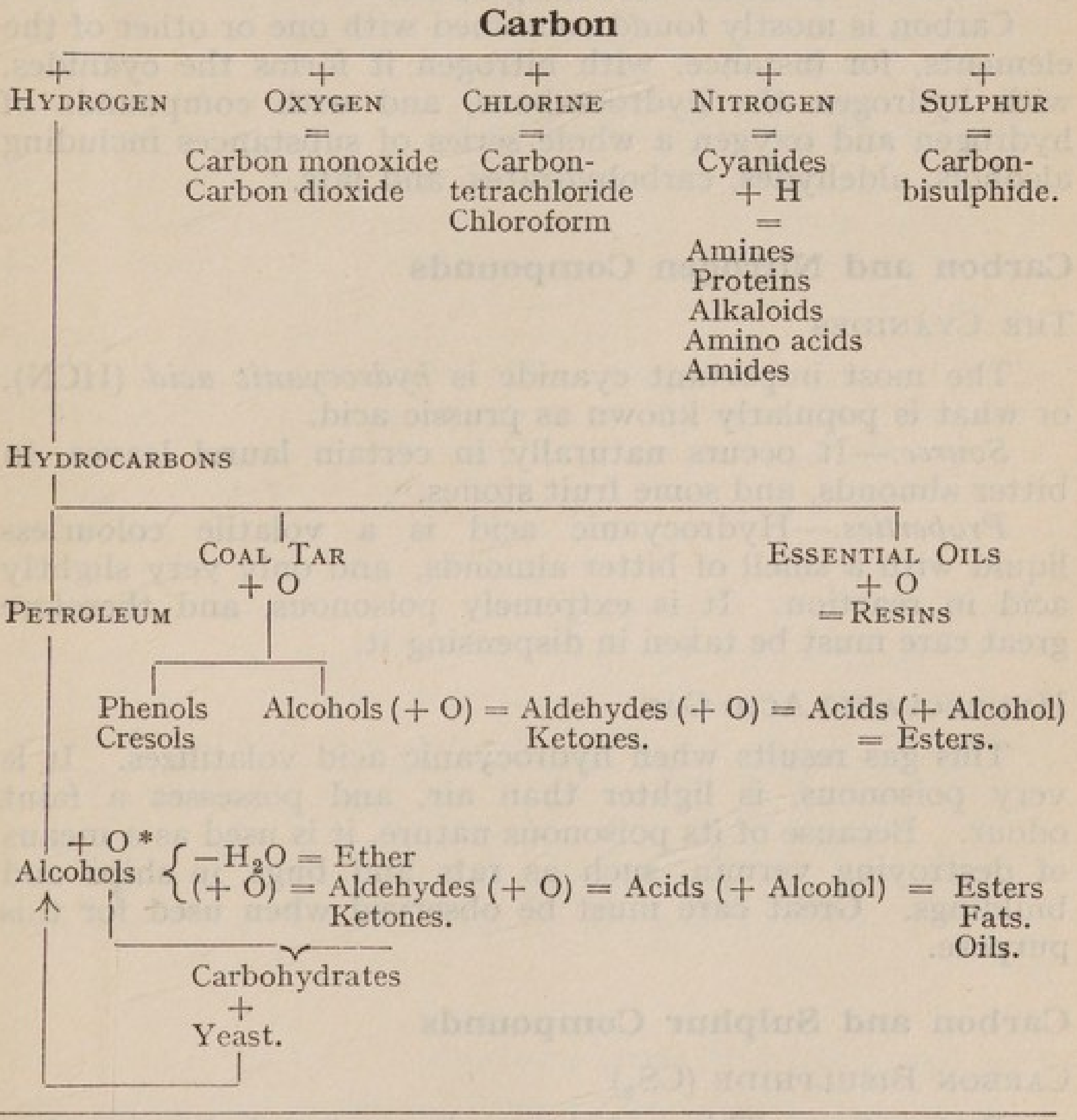
ORGANIC CHEMISTRY

CHAPTER 1

ALL substances that enter into the composition of, or are prepared from, animal or vegetable matter are called organic substances.

These organic substances always contain carbon, and for this reason are sometimes known as the Carbon Compounds, and they comprise an immense number of substances.

TABLE SHOWING RELATIONSHIP OF ORGANIC CARBON COMPOUNDS



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* The alcohols are always produced from carbohydrates by fermentation as oxidation of the petroleum oils is difficult.

In this book they have been given very little space, and it must not be understood by this that they are unimportant. To the dispenser they are in fact as important or more so than the inorganic substances, but as Materia Medica deals with them in detail, the intention is only to give the student a general idea of their classification and grouping so that he can the better appreciate Materia Medica when he comes to study it.

In the accompanying table an attempt has been made to show the relationship between the various groups of organic substances, but it must not be thought that because a substance comes directly under another substance in the table it is manufactured from it. This is definitely not so in the majority of cases, for example, alcohol comes under the petroleum hydrocarbons, but it is not produced from any of these, but from the fermentation of sugar, etc.

Carbon is mostly found combined with one or other of the elements, for instance, with nitrogen it forms the cyanides, with hydrogen the hydrocarbons, and with compounds of hydrogen and oxygen a whole series of substances including alcohols, aldehydes, carbohydrates, and fats.

Carbon and Nitrogen Compounds

THE CYANIDES

The most important cyanide is *hydrocyanic acid* (HCN), or what is popularly known as prussic acid.

Source.—It occurs naturally in certain laurel leaves, in bitter almonds, and some fruit stones.

Properties.—Hydrocyanic acid is a volatile colourless liquid with a smell of bitter almonds, and only very slightly acid in reaction. It is extremely poisonous, and therefore great care must be taken in dispensing it.

HYDROCYANIC ACID GAS

This gas results when hydrocyanic acid volatilizes. It is very poisonous, is lighter than air, and possesses a faint odour. Because of its poisonous nature, it is used as a means of destroying vermin, such as rats and bugs, in ships and buildings. Great care must be observed when used for this purpose.

Carbon and Sulphur Compounds

CARBON BISULPHIDE (CS_2)

Preparation.—This substance is prepared by burning carbon in sulphur vapour.

Properties.—It is a colourless, volatile and very inflammable liquid which has no smell when fresh, but develops a disagreeable and characteristic odour when stale. It is a poisonous substance, and is used for the purpose of killing insects. It is also a solvent of certain substances such as rubber.

QUESTION

Question.—Into what groups can the organic carbon compounds be divided ? Give examples of each group.

CHAPTER 2

CARBON AND HYDROCARBON COMPOUNDS

The Hydrocarbons

THERE are many substances belonging to this group, a considerable number of which are entirely different in appearance, yet they all contain only carbon and hydrogen, the difference depending on the proportions of carbon and hydrogen present.

All the hydrocarbons can be included in three main groups, namely, the Petroleum Hydrocarbons, the Coal Tar Hydrocarbons, and the Essential Oil Hydrocarbons.

The petroleum hydrocarbons have no chemical action except combustion.

The Petroleum Hydrocarbons

The origin of this group of hydrocarbons is from oil wells in various parts of the world, where it emerges as a dark greenish-brown thick liquid, and from it, by a process of distillation, the following substances are separated.

The more volatile substances distil over first. Amongst these are petroleum ether, aviation spirit, motor spirit, and paraffin oil, all highly inflammable substances. Then come the heavier oils, light and heavy lubricating oils, and finally the solid residues which when purified give vaseline and paraffin wax.

The Coal Tar Hydrocarbons

The origin of this group is in every case from coal, generally by the destructive distillation of this substance. A crude tar is the first result, and when this has been treated chemically to remove carbolic acid which is present, the following substances can be obtained by distillation in a similar manner to that in the preparation of the petroleum group: benzene or benzol and xylol, which are highly inflammable fluids, naphthalene or moth balls, a white crystalline solid, and the residue, pitch or coal tar.

The Essential Oil Hydrocarbons

These oils give rise to the scent of plants from which they are extracted. Some of the more common are sold as perfumes, such as lavender oil, attar of roses, oil of thyme, etc.

Other essential oils are used for flavouring, for example, oil of lemon, while others are used as deterrents to repel insects, such as oil of cassia, and oil of citronella.

Turpentine is an example of an essential oil which is used in medicine.

PRACTICAL CHEMISTRY—CHAPTER 2

Expt. 135

The Petroleum Hydrocarbons

Demonstrate that the petroleum hydrocarbons have no chemical action by treating liquid paraffin with strong sulphuric acid in a test-tube. There is no charring.

QUESTION

Question.—Into what groups are the hydrocarbons classified ? Give examples of each group, and state briefly how they are prepared.

CHAPTER 3

ALCOHOLS AND PHENOLS

In the composition of many organic substances there are present certain combinations of elements which do not break up when chemically attacked, and act more or less like metallic elements. Such combinations are known as radicals, and are similar to the compound radical NH_3 (Ammonia), which remains unaffected in itself by chemical action.

Hydroxides of the Petroleum Hydrocarbons

Any one of the petroleum hydrocarbons may act as a radical and form hydroxides which are called alcohols. In theory alcohols can be produced in this manner, but actually they are prepared in the method described under each.

METHYL ALCOHOL OR WOOD SPIRIT is produced by the distillation of wood. It is a colourless liquid, lighter than water, very mobile and inflammable, burning with a non-luminous flame.

ETHYL ALCOHOL OR SPIRIT occurs when sugar is fermented by yeast. By further treatment, pure or absolute alcohol can be obtained, while there are other grades of alcohol, such as "rectified spirits," "industrial spirit," and "surgical spirit," which are mixtures of ethyl alcohol and various other substances.

ISO-PROPYL ALCOHOL is another alcohol belonging to the petroleum series. It is less commonly known, yet it is used sometimes in the preparation of tinctures.

Hydroxides of the Coal Tar Hydrocarbons

The coal tar hydrocarbons can also act as radicals, and form phenol, cresol, etc., when (OH) is linked to them.

PHENOL is a hydroxide of one of the coal tar hydrocarbons. It is not an alcohol, and in the impure state is known as carbolic acid. Similarly, *Cresol* is an hydroxide of another of the coal tar hydrocarbons.

They both occur in coal tar, from which they are extracted by treatment with an alkaline solution followed by distillation.

In the pure state they are solid, but the crystals dissolve rapidly on heating, or on the addition of a little water. They have a distinctive carbolic odour and are poisonous, being used extensively as antiseptics and disinfectants. In the latter case, when combined with soap they form the emulsion—liquor cresoli saponatus fortis. Care must be taken when handling concentrated phenol or carbolic acid, as it blisters or burns the skin.

PRACTICAL CHEMISTRY—CHAPTER 3

Expt. 136

Alcohols

Into an inclined test-tube containing a little water pour gently some alcohol; the alcohol floats on the top of the water.

Shake and an even mixture results.

Expt. 137

Phenol or Carbolic Acid

Place some crystals of carbolic acid in a test-tube taking great care not to allow them to touch the fingers.

Heat very gently; the crystals soon melt.

Cool by placing the end of the test-tube in cold water; the crystals reform. Add two or three drops of water; they dissolve rapidly.

QUESTION

Question.—What is ethyl alcohol, and how is it prepared? Name its more important properties.

CHAPTER 4

ALDEHYDES, KETONES AND ORGANIC ACIDS

Aldehydes

When alcohol is oxidized aldehydes are generally produced; thus methyl alcohol becomes formaldehyde gas, and ethyl alcohol becomes the liquid acetaldehyde.

FORMALDEHYDE GAS is a pungent irritating gas which is soluble in water, a saturated solution (40 per cent.) being sold as *formalin.*

On standing or heating this formalin solution, a white powder, *paraform,* settles, which when dried and heated gives off formaldehyde gas. It is from formalin and paraform tablets that formaldehyde gas is generated for disinfecting purposes.

When the gas is bubbled through ammonia, crystals of *hexamine* or *urotropine* are formed. This substance is a urinary disinfectant, because in the body it breaks up again into ammonia and formaldehyde gas, which latter acts as an internal disinfectant.

ACETALDEHYDE.—This is the oxidized product of ethyl alcohol (rectified spirit), and appears as a clear volatile liquid, soluble in water, and with a sweet fruity smell.

When strong sulphuric acid is added to acetaldehyde in the proper proportions a liquid is derived which is known as *paraldehyde,* a drug used as a hypnotic or sleep producer.

CHLORAL HYDRATE may be mentioned here as a product resulting from the mixing of chlorine gas and alcohol with some further chemical treatment and final distillation.

An oily liquid, known as *chloral,* distils over, which when moistened with water is transformed into the white crystals of chloral hydrate.

Chloral hydrate is used to produce sleep.

Ketones

Aldehydes do not invariably result from the oxidation of alcohol, but substances somewhat similar to aldehydes, known as *ketones,* are produced. *Acetone* is the best known of the ketones, and is produced by the oxidation of iso-propyl alcohol, and can be detected by its characteristic sweet fruity odour.

When phenols and cresols are oxidized they do not produce aldehydes, but complex mixtures or compounds.

Organic Acids

Organic acids occur widely distributed in nature, in animals and in plants, or can be isolated from fats and oils; for example, formic acid occurs in the stings of bees and wasps, also of nettles; acetic acid is present in vinegar which is made from beer by the action of certain germs or by the distillation of wood. Oleic, palmitic, and stearic acids can be separated from their respective fats and oils.

Citric acid gives the tart taste to lemons, oranges, and limes; salts of oxalic acid are found in rhubarb, while tartaric acid can be extracted from one of its salts present in wine.

These are only a few of the best known organic acids, which can be divided into two main classes—the *petroleum acids* or *fatty acids*, which include all those mentioned above, and the *coal tar acids*, of which the two in common use in pharmacy are benzoic and salicylic acids.

The Petroleum Acids or Fatty Acids

To repeat, these include, among many others, *formic*, *acetic*, *oleic*, *palmitic*, *stearic*, *tartaric*, and *citric* acids. As well as occurring in nature they can be prepared artificially, and this is the method when manufactured on a large scale.

The different acids can be isolated from their salts by the addition of a mineral acid; thus sulphuric acid added to sodium acetate gives acetic acid.

Some appear as solid, others as liquid substances; for example, tartaric and citric acids are crystalline solids, palmitic and stearic are waxy solids, while oleic, formic and acetic acids are liquid. They are all weak acids, and for this reason are easily displaced from their salts by mineral acids. They are all soluble in alcohol and, with the exception of oleic, palmitic, and stearic, also in water. These organic acids unite with alcohol to form esters.

Coal Tar Acids

There are many acids in this series, but the two of outstanding importance to the pharmacist are *benzoic* and *salicylic acid*.

Benzoic acid occurs naturally in gum benzoin, in the balsams of Peru and Tolu, while *salicylic acid* occurs in oil of wintergreen, but both these acids when prepared on a large scale are obtained from coal tar.

They are white crystalline substances nearly insoluble in water, but soluble in alcohol and ether. They form salts with bases, and esters with alcohols in a similar manner to the fatty acids.

Expt. 138

Aldehydes.—Formaldehyde

In a test-tube place a few crystals of potassium permanganate and moisten them with 40 per cent. formalin. After a short interval pungent fumes of formaldehyde gas are given off.

Expt. 139

Acetaldehyde

Pour a tea-spoonful of methylated spirits into a test-tube, add a salt-spoonful of chromic acid and warm cautiously. A sweet fruity smell of acetaldehyde comes off.

Expt. 140

Organic Acids

Place a drop of acetic acid or vinegar, or a moist crystal of salicylic acid, on a piece of blue litmus paper; the paper turns red, indicating that these substances are acid in reaction.

Expt. 141

Organic Acids are Weak Acids

Boil a little lead acetate with dilute hydrochloric acid in a test-tube; fumes of acetic acid are given off, and can be detected by their odour.

QUESTIONS

Question 1.—What are aldehydes? Name a common example, stating its properties.

Question 2.—Name a few examples of organic acids, stating the source of each. How are organic acids classified?

CHAPTER 5

THE CARBOHYDRATES—SUGARS, STARCHES, ETC.

CARBOHYDRATES are abundantly found in the vegetable and animal kingdom. Those from a vegetable source are: *cane* sugar, which is manufactured from sugar cane or from beet; *glucose* or *grape* sugar which is present in the grape; *starch* which is especially plentiful in tubers and cereals; and cellulose in wood fibre. The carbohydrates from an animal source include *glycogen*, which is found mainly in the livers of animals, and *lactose* or *milk sugar*, present in milk.

These substances contain carbon, hydrogen, and oxygen, the molecules of carbon are six or more, and the hydrogen and oxygen are in the proportion of two to one as in water. Carbohydrates can be classified as monosaccharides with a formula $C_6H_{12}O_6$, disaccharides $C_{12}H_{22}O_{11}$, and polysaccharides $C_6H_{10}O_5$.

The mono- and disaccharides are all crystalline and soluble substances, while the polysaccharides are mainly colloidal, and therefore generally insoluble.

Monosaccharides ($C_6H_{12}O_6$)

The commonest sugar of this series is *glucose* or grape sugar. It is only in the form of glucose that sugars can be absorbed by and circulate in the body, and therefore all disaccharides and polysaccharides must be reduced to glucose before absorption can take place. This reduction is effected by certain digestive juices in the alimentary canal.

Monosaccharides when acted on by yeast, ferment and form alcohol and carbon dioxide.

Disaccharides ($C_{12}H_{22}O_{11}$)

Cane sugar is of all sugars the best known, as it is the one which is used daily on the tea table. *Lactose* or milk sugar is present in the milk of all mammals, and *maltose* or malt sugar is present in grain when it sprouts.

These sugars ferment with yeast in the same manner as those of the preceding series.

Polysaccharides ($C_6H_{10}O_5$)

These are all chiefly colloidal substances which are insoluble in water, and those that can be broken down by the process of digestion are reduced in the body to disaccharides, and then to monosaccharides before absorption.

Of this series, starch, glycogen, inulin, and cellulose are the commonest members. *Starch* is the sugar stored in plants, while *glycogen* is the sugar stored in animals, mainly in the liver. In both cases, for the uses of the plant or the animal, the starch or glycogen must be converted into soluble sugar before it can circulate in the sap or in the blood. Thus it is cane sugar that is found in the sap of plants, and glucose in the blood of animals.

Starch occurs in all cereals, legumes (peas and beans), and tubers such as potatoes, turnips, etc. The little grains of starch are surrounded by a cellulose envelope which ruptures when the grains are boiled, liberating the gelatinous starch. The presence of starch can be at once detected by the addition of iodine, when the brown colour of the latter is changed to blue.

Inulin is a starchy substance derived from the root of the dahlia plant and certain other roots and tubers, and is used in the treatment of diabetes. It should not be confused with insulin, which is prepared from the pancreas of certain animals.

Cellulose constitutes the chief material, comprising the framework of plants; for example, wood fibre and straw are mainly composed of cellulose. It is not digested by human beings, and is excreted from the body unaltered.

PRACTICAL CHEMISTRY—CHAPTER 5

EXPT. 142

Carbohydrates

To show they are Carbon Compounds, burn some sugar or starch in a tin lid, and black carbon will be left.

EXPT. 143

Solubility of Mono- and Disaccharides, insolubility of Polysaccharides

Shake up in three test-tubes of water, (*a*) glucose, (*b*) cane sugar, and (*c*) starch, in small quantities. Glucose and cane sugar dissolve and starch remains insoluble.

EXPT. 144

Starch a Colloid

Shake up the starch in the tube and boil—a colloidal solution is formed. Add some tincture of iodine, a blue colour results.

QUESTION

QUESTION.—State what you know of the composition, grouping and properties of carbohydrates.

CHAPTER 6

FATS, OILS, AND SOAPS

FATS and oils may be divided into two classes, those that form soaps with hot alkalis and those that do not. Of the former *stearin*, *palmitin* and *olein* are the best known, and of the latter such oils as paraffin, light and heavy lubricating oil, which are mentioned under the petroleum hydrocarbons, pitch and coal tar of the coal tar hydrocarbons, and any of the essential oils.

Fats and Oils

STEARIN, PALMITIN, AND OLEIN

The true fats and oils are compounds of fatty acids, such as stearic, palmitic and oleic, with glycerine, which belongs to the alcohols, although in character not similar. As already stated, fatty acids and alcohol give rise to esters, so that fats and oils are really esters.

Stearin is a hard fat like suet, with a high melting point; palmitin is a soft fat like lard, which melts at a low temperature; while olein is an oil which remains fluid at normal temperature. Examples of oils are many—cocoanut oil, olive oil, linseed oil, cod-liver oil, etc.

Properties.—Fats and oils are insoluble in water and in order to get a thorough mixing of the two an emulsion must be formed. They are, however, soluble in ether, benzine, petrol, etc., which substances are commonly used for the removal of greasy stains from clothing. They all burn when heated, and if allowed to become rancid split up into glycerine and their respective fatty acids which give rise to the unpleasant flavour. They all form soap when boiled with alkalis.

Soaps

These are prepared by mixing the melted fat or oil with a boiling solution of caustic soda or caustic potash, adding salt which causes the soap to rise to the top, by which means it can be removed and collected. It is further purified and perfumes or antiseptics are added if desired. The residue contains glycerine which is a very valuable by-product. Thus,

$$\text{Stearin (fat)} + \text{alkali (caustic soda)} = \text{Glycerin} + \text{Sodium stearate (soap)}.$$

If the alkali used is caustic soda a hard soap is produced, whereas caustic potash gives rise to a soft soap. Soaps that contain an excess of alkali are harmful to delicate skins, and therefore manufacturers are very careful in removing every trace of it.

Properties.—Soaps give a frothy solution with water and a clear solution with alcohol. With a hard water, that is one containing calcium and magnesium salts, it takes a long time before a lather with soap can be formed; in fact it is not until all these salts are removed and the water softened that a permanent lather results.

Use has been made of this fact in determining the hardness of a water.

PRACTICAL CHEMISTRY—CHAPTER 6

EXPT. 145

Fats and Oils

Soap-forming and Petroleum Oils

Place a drop of liquid paraffin and a drop of olive oil into two separate test-tubes, pour a teaspoonful of glycerine into each and half a teaspoonful of caustic soda solution, and heat until white fumes come off. Pour into a glass of cold water.

In the case of liquid paraffin the oil drops will rise to the surface unchanged, but in the case of olive oil a soap solution will be formed, which will make a froth if the water is not too hard.

CHAPTER 7

AMINES, AMIDES AND AMINO ACIDS, ALKALOIDS

THESE three groups of substances contain nitrogen, combined with carbon, hydrogen, and, with the exception of the amines, with oxygen.

The Amines

These are mainly oily evil-smelling substances, found in the products of decomposed flesh, and sometimes termed ptomaines.

The Amides

These include *urea*, which is present in urine.

The Amino Acids

When meat and egg albumin, which are called proteins, are treated with acid, either artificially or naturally as in the digestive system, they are converted into white crystalline substances called amino acids.

The protein molecule is composed of many different amino acids, but although all can be prepared in the laboratory, chemists have failed to make them combine to reform protein.

Alkaloids

The alkaloids are obtained from plants; they contain nitrogen and are very similar to the amines.

They have all a very powerful effect on man, even when administered in the smallest doses, and in some cases are extremely poisonous.

Alkaloids are mainly solid white substances, having an intensely bitter taste and are almost insoluble in water, although their salts are easily dissolved. Those commonly used in medicine are *morphine* from opium, *quinine* from cinchona bark, *strychnine* from the seeds of nux vomica, *atropine* from the deadly nightshade or belladonna plant, and *cocaine* from the coca plant.

OPIUM AND MORPHINE

When the juices from certain poppy seeds are dried a brown amorphous substance called opium is obtained which, in addition to many other alkaloids, contains the well-known substance—morphine.

Medicinally opium is used to promote sleep or allay pain, its effects mainly being due to the morphine present.

Morphine is a white crystalline powder, practically insoluble in water, ether and chloroform, but its salts, for example, the hydrochloride, are soluble.

It is an extremely poisonous substance, producing unconsciousness in large doses, and is used medicinally for the same purposes as opium.

QUININE

This alkaloid, along with many others, is extracted from the cinchona tree. It is a white crystalline substance, almost insoluble in water but more soluble in ammonia solutions and in ether and chloroform.

It is not so poisonous as some of the other alkaloids, but possesses their characteristic bitter taste.

STRYCHNINE

Strychnine occurs in the seeds of the nux vomica plant; and both it and its salts can be identified by their characteristic appearance.

The salts are all soluble in water and alcohol, but strychnine is insoluble in both these and in ether, although soluble in chloroform. It is intensely poisonous and is given medicinally in the smallest doses.

ATROPINE

Atropine is obtained from the deadly nightshade or belladonna plant.

Although itself only slightly soluble its salts dissolve readily in water. It is used a great deal in the treatment of eye conditions as its action is to dilate the pupil.

COCAINE

This alkaloid is obtained from the coca (not cocoa) plant.

It is a powdery substance which like other alkaloids is feebly soluble in water, while its salts are easily dissolved. It is very poisonous, both when taken by the mouth and injected as a solution, when it acts as a local anæsthetic, dulling all sensation in the neighbourhood.

In tea and coffee there are small amounts of the relatively harmless alkaloids—caffeine and theine, while in cocoa theobromine is present.

Nicotine is an alkaloid present in the tobacco plant, and sometimes gives rise to symptoms in persons who are heavy smokers.

PRACTICAL CHEMISTRY—CHAPTER 7

Alkaloids

Expt. 146

Solubility and Taste

Shake up some quinine and some quinine bisulphate with water in separate test-tubes; quinine does not dissolve, but quinine bisulphate does eventually.

Wet the finger with this solution and place on the tip of the tongue; an intensely bitter taste characteristic of all alkaloids will be noticed.

QUESTION

Question.—Describe two examples of alkaloidal substances and state the main properties of each.

CHAPTER 8

CHLORINE COMPOUNDS OF CARBON

ANÆSTHETICS

THESE compounds contain carbon, hydrogen and chlorine, the characteristic properties of the latter being entirely absent. They are all liquids and do not exist in nature, being manufactured by various chemical methods.

Those that are best known to the dispenser are methyl chloride, chloroform, and carbon tetrachloride.

Methyl Chloride

It can be prepared by mixing the gas methane, one of the petroleum hydrocarbons, with chlorine gas. It is a volatile liquid which evaporates rapidly when exposed to the air, and in so doing produces intense cold. It is used as a local anæsthetic, freezing the skin, thus dulling the pain of a knife incision.

Chloroform

It can be prepared in various ways, but the usual method is by distilling a mixture of alcohol, bleaching powder and water. Chloroform is dealt with in greater detail under anæsthetics.

Carbon Tetrachloride

Carbon tetrachloride is prepared from carbon disulphide and chlorine. It is very similar to chloroform, but has not such good anæsthetic properties. It is non-inflammable and can be used for extinguishing burning petrol, ether, etc. It is toxic to insects and is the main constituent in many fly sprays.

Iodine Compound of Carbon

Iodoform

Iodoform is prepared from alcohol, iodine and soda, and appears as a yellow amorphous powder with a powerful and persistent odour.

It is used for antiseptic and disinfectant purposes.

Anæsthetics

The term anæsthetic is applied to a substance which produces either total or local insensibility to pain. When complete unconsciousness results the anæsthetic is known as a general anæsthetic ; when the insensibility is localized to the area injected, then the anæsthetic is termed a local anæsthetic.

General Anæsthetics

General anæsthetics include the extremely volatile liquids—ether, chloroform, and ethyl chloride, or the gas nitrous oxide, which has been described in the inorganic section.

ETHER

Ether is produced by dropping alcohol on hot sulphuric acid, contained in a flask and distilling over the ether formed.

There are certain impurities, such as alcohol and water, that are first removed, and then by a process of rectification the *anæsthetic ether* is obtained.

Properties.—Ether is a clear mobile fluid which volatilizes rapidly, and is very inflammable. In the administration of ether the fluid is allowed to evaporate, and the ether vapour is inhaled producing anæsthesia. Ether should be kept in a yellow-coloured bottle to prevent deterioration, as it slowly oxidizes when not properly stored and undesirable and dangerous products are formed.

CHLOROFORM

Preparation.—Chloroform is prepared from alcohol, bleaching powder and water, by a process of distillation. The distillate is crude chloroform, which is further treated by rectification to obtain the pure substance.

Properties.—It is a clear mobile volatile liquid ; the method of its administration and its action in producing unconsciousness is similar to that of ether.

It must be stored in the dark and in coloured bottles, otherwise it decomposes, and products appear which are irritating to the lungs.

Local Anæsthesia

The local anæsthetics most commonly in use are : *cocaine, eucaine, novocaine,* and *stovaine.*

Cocaine has already been described when dealing with the alkaloid. In medicine it is the hydrochloride of cocaine that is used because it is soluble.

Eucaine, novocaine, and stovaine are chemically prepared substances with complex formulæ, and which are all soluble. When dissolved in water they are either injected under the skin, into the spinal canal, or painted on the surface to be anæsthetized.

There are certain properties which a local anæsthetic must possess :

(1) It should dissolve in water.
(2) It should not set up inflammation or irritation.
(3) It should not lose its anæsthetic power on sterilizing.

Methyl chloride, by freezing the part and dulling sensation, is used as a local anæsthetic.

Gums, Mucilages, and Resins

These substances are all products from plants, and are somewhat similar in their physical properties, being sticky substances.

Gums

These are contained in certain plants, and are carbohydrate in nature, taking a position as regards their properties between the sugars and the starches.

They are soluble in water like the sugars, but they make a very much more sticky solution, and they are used in dispensing to keep insoluble substances in suspension, preventing them from settling.

The gums are precipitated from their solution in water by alcohol ; therefore alcohol should not be added to a mixture containing gum. Gum arabic is the ordinary gum of commerce, gum acacia and gum tragacanth are those used in dispensing. Although gum tragacanth is included, it is not strictly speaking a gum, but a mucilage.

Mucilages

Very little is known of these substances ; when mixed with water they swell up and make a slimy mixture, but do not dissolve. Tragacanth is an example of a natural mucilage occurring in plants.

Resins

Resins are substances which are obtained from plants ; they originate from the essential oils in the sap of certain trees, such as the pine, and are formed from these oils by exposure to air. Thus oil of turpentine is the essential oil of the pine, and when the bark is bruised, the turpentine will exude and on exposure to the air will oxidize and harden into resin.

Many other plants, shrubs and trees have their own special resins, for example, scammony, jalap, etc., and these have their particular action on the body—jalap, for instance, is a purgative.

The resins are very weak acids and are insoluble in water, but soluble in alcohol; thus they are very different substances from gums and mucilages.

PRACTICAL CHEMISTRY—CHAPTER 8

Expt. 147

Iodoform

Mix equal portions of tincture of iodine and caustic soda solution (10 per cent.) in a test-tube; allow it to stand, when a yellow precipitate and a strong odour of iodoform will be produced.

Expt. 148

Ether. Inflammability of Ether and Chloroform

(i) Pour a drop of ether in a tin lid, remove the bottle of ether to a safe distance and place a lighted match to the ether; it catches fire at once and burns with a non-luminous flame.

(ii) Repeat experiment with chloroform; the match will be put out.

Expt. 149

Properties of Ether and Chloroform

Repeat Expt. 135, using ether and chloroform in two separate tubes.

Ether rises to the surface of the water and does not form a fluid mixture; chloroform sinks without mixing.

Ether is therefore lighter and chloroform heavier than water.

Expt. 150

Gums

Dissolve some gum arabic in water in a test-tube and mix with an equal volume of methylated spirits. The solution becomes milky owing to the precipitation of gum.

Expt. 151

Mucilage

Add some tragacanth powder to some warm water in a test-tube, it slowly swells without dissolving; this distinguishes it from gums, which dissolve as shown in the experiment above.

DEFINITIONS

Acid. A compound containing one or more atoms of hydrogen, replaceable by a metal, having a sour taste, and turning blue litmus red.

Accumulator. An apparatus for storing electricity.

Affinity. The force that unites atoms into molecules.

Alkali. A compound which forms soluble soaps with fatty acids, and turns red litmus blue.

Allotropy. The existence of an element in two or more distinct forms (allotropic forms) with distinct physical properties.

Alternating current. A current which periodically flows in opposite directions.

Ampere. The unit that measures the flow of an electric current.

Amorphous. A solid substance having no definite shape.

Analysis. The breaking down of a substance into its component parts.

Anhydrous. Without water of crystallization.

Anhydride. A chemical compound derived from a substance, especially an acid, by the abstraction of a molecule of water.

Anode. The positive pole of a galvanic battery.

Antiseptic. A substance that will prevent the growth of germs without destroying them.

Atmospheric pressure. The weight of the atmosphere, which at sea-level is 15 lb. to the square inch.

Atom. The smallest division of matter which can take part in a chemical change.

Base. The non-acid part of a salt ; a substance which combines with acids to form salts.

Bleaching agent. A substance which removes colouring matter.

Boyle's Law. The volume of a gas varies inversely as the pressure, the temperature being constant.

Calorie. A unit of heat, being the amount of heat required to raise 1 gramme of water 1 degree centigrade.

Capillary attraction. The force which attracts the particles of fluid into and along the bore of a narrow tube.

Cathode. The negative pole of a galvanic battery.

Charles's Law. On being heated all gases expand by the same amount, the pressure being constant.

Chemical change. A complete change in the composition of a molecule.

Chemically active substance. A substance which readily alters its composition by the action of chemicals or of heat.

Chemically inactive substance. A substance which remains unaltered in spite of the action of the strongest chemical agents.

Chemistry. Is the science which treats of the composition and properties of all matter.

Circuit. The round or course traversed by an electric current.

Clarification. The clearing from a liquid of its turbidity.

Cohesion. The force which unites the particles of a body.

Colloids. Substances whose particles are big enough to remain evenly distributed throughout the fluid and yet not big enough to settle down by the action of gravity.

Combustion. Burning or rapid oxidation with emission of heat.

Combustible material. Matter that can be burned.

Compound, A. A substance which consists of two or more chemical elements in union.

Concave. Presenting a hollow surface.

Condensation. The process of conversion of gases and vapours into fluid.

Condenser. A vessel or apparatus for condensing gases or vapours.

Conductor of electricity. A substance that allows electricity to flow through it.

Conductor of heat. A substance that allows heat to flow through it.

Convection of heat. The carriage of heat by currents of air.

Convex. Presenting a bulging surface.

Crystalline substance. A solid substance which assumes a definite geometrical shape.

Crystalloid. A substance which in solution passes readily through animal membranes.

Current electricity. That type of electricity produced by chemical or magnetic means and which flows along a wire.

Decomposition. The breaking down of chemical substances into simpler bodies.

Deliquescence. The condition of becoming liquefied as a result of the absorption of water from the air.

Density. The quantity of matter in a given space.

Deodorant. A substance which removes undesirable or offensive odours.

Desiccation. The act of drying up.

Desiccator. A closed vessel for containing chemicals that are to be kept free from moisture.

Diffusion. The process of becoming widely spread.

Direct current. An electric current which flows in one direction only.

Dialysis

Disinfectant.	An agent that kills germs.
Distillation.	The process of converting a substance into a gaseous or vapour form by heat, and then cooling to a liquid and collecting the evaporated matter.
Distillate.	The portion that has been distilled over.
Divalent.	Capable of uniting with or displacing two atoms of hydrogen.
Ductibility.	Capable of being drawn out, as into a wire.
Dynamo.	A machine for converting mechanical force directly into current electricity.
Efflorescence.	The state of becoming powdery in consequence of losing the water of crystallization.
Electrolyte.	Any substance which in solution conducts electricity, and is decomposed by it.
Element.	A simple substance which cannot be decomposed by chemical means and which is made up of similar atoms.
Electrolysis.	Chemical decomposition produced by passing a direct current through the solution of a compound.
Electro magnet.	A temporary magnet made by passing an electric current through a coil of wire surrounding a piece of soft iron.
Emulsion.	A mixture of very minute drops of oil with water.
Equation (*chemical*).	An equation which expresses a chemical reaction ; the symbols on the left of the equation denoting the substances before and those on the right after the reaction.
Evaporation.	The conversion of a solid or a liquid into vapour.
Fermentation.	A chemical or physical change brought about by a ferment such as yeast.
Filter.	A device for straining water or other liquid.
Filtrate.	The liquid which has passed through a filter.
Filtration.	The passage of a liquid through a strainer which keeps back insoluble particles.
Formula.	A combination of symbols to express the chemical composition of a substance.
Gas.	A substance capable of expanding indefinitely into space.
Hydrate.	Any compound of a radical with water.
Hydrometer.	An instrument for finding the specific gravity of fluids.
Hypnotic.	A drug that induces sleep.

Indicator. Any substance which, when added in small quantities, shows the appearance or disappearance of a substance by a conspicuous change of colour.

Insulation. The stoppage of an electric current or heat by a non-conducting substance.

Insulators. An appliance made of some non-conducting material used in preventing the escape of electricity or heat.

Inflammable. Capable of being set on fire.

Latent heat. The heat that a body absorbs when changing from one physical state to another without any alteration in its temperature.

Liquefaction. A change to liquid form.

Liquid. A substance which assumes the shape of the containing vessel and has a level surface.

Luminosity. The property of giving off light.

Magnet. A piece of lodestone, or a bar of steel or iron that has the power of attracting iron.

Magnetic field. The area in the vicinity of a magnet in which the attraction for iron exists.

Malleability. Capable of being beaten out into a thin plate.

Matter. Anything that occupies space.

Meniscus. The surface of a column of liquid.

Metal. Any element which possesses lustre, malleability, ductility, and conductivity of electricity and heat.

Metalloid. Any metallic element which has not all the characteristics of a typical metal.

Molecule. The smallest particle of a substance which can exist by itself and which is composed of a collection of atoms.

Monovalent. Capable of uniting with or displacing one atom of hydrogen.

Negative electricity. That type of static electricity which is produced by rubbing resin, amber, or sealing wax with flannel.

Neutral. Neither acid nor alkaline.

Neutralize. To render a solution neither acid nor alkaline.

Non-metal. Any chemical element which is not a metal or metalloid.

Non-stable. Easily decomposed by physical or chemical means.

Ohm. The unit of resistance to an electrical current.

Opaque. The property of preventing the passage of rays of light.

Osmosis. The passage of a liquid through a membrane.

Osmotic pressure. The pressure which brings about diffusion between solutions of different concentration.

Oxide. A compound of oxygen with an element or radical.

Oxidation. The process by which atoms of oxygen are chemically united to a substance.

Oxidizing agent. A substance possessing the power of giving up its oxygen readily.

Physical. Pertaining to nature.

Physical change. A change in the nature or appearance of a substance without any alteration in its chemical composition.

Plane. Presenting a level surface.

Polarization. The accumulation of bubbles of hydrogen gas on the negative plate of a galvanic battery.

Positive electricity. That type of static electricity which is produced by rubbing glass with silk.

Potential. The power or pressure of electricity a body possesses.

Precipitation. The formation of solid particles in a liquid, the result of mixing two solutions.

Radiant heat. Heat that passes as waves through the air, heating only those objects which it hits.

Rancidity. A term applied to fats that have decomposed with the formation of fatty acids.

Reaction. A chemical process in which one substance is transformed into another substance or substances.

Rectification. The re-distillation of a liquid to purify it.

Reducing agent. A substance possessing a great affinity for oxygen and having the power of removing oxygen from another substance.

Reduction. The removal of oxygen from a substance.

Reflection of light. The turning back of a ray of light when it strikes against a surface which it will not penetrate.

Refraction. The bending of light on passing from one medium to another of different density.

Resistance. The opposition by a substance to the passage of an electric current.

Salts. Any compound of a base or radical and an acid.

Solid. A substance which retains its shape.

Solution. A liquid containing some other substance dissolved in it.

Solvent. A liquid that is capable of dissolving substances.

Specific gravity. The weight of a substance compared with that of an equal volume of water.

Spectrum. A band of different colours into which light is decomposed on passing through a prism of glass.

Stable. Not easily decomposed by physical or chemical means.

Static electricity. That type of electricity produced by friction and which remains stationary in the substance.

Sublimate. A substance prepared by sublimation.

Sublimation. The process of vaporizing and condensing a solid substance without melting it.

Surface tension. The strain or resistance which prevents the surface of a liquid rupturing.

Suspension. A mixture of an insoluble solid and a liquid where the solid particles are distributed equally throughout the liquid and held there by the addition of some mucilaginous substance.

Symbol. A letter or combination of letters representing an atom of an element or a molecule of a radical.

Tarnish. The coating of the metallic oxide found on metals when exposed to the air and obscuring the metallic lustre.

Translucency. The property of allowing rays of light to pass so that objects are not seen clearly through the substance.

Transparency. The property of allowing light to pass so that objects are seen clearly through the substance.

Trivalent. Capable of uniting with or displacing three atoms of hydrogen.

Valency. The power possessed by an element or radical of uniting with or displacing hydrogen in a compound.

Viscosity. The property of being sticky or gummy.

Volatile. Tending to evaporate rapidly.

Volt. The unit of electric pressure or force.

Water of crystallization. The water which is contained in many salts, giving them a crystalline form.

Weight. The degree to which a body is drawn toward the earth by gravity.

APPENDIX 1

APPARATUS AND CHEMICALS REQUIRED

Apparatus

Description	Approx. quantity for class of six	Suggested source
Aluminium	—	Pieces of old cooking vessel.
Biscuit tin, empty	No. 1	Ration store.
Black lead	oz. ½	—
Bottle, ether, empty	No. 1	Dispensary.
,, medicine	,, 1	,,
Beakers, sizes 1 to 4	,, 4	Special supply or district laboratory.
Candles	,, 1	Dispensary.
Corks, pint and quart	,, 6	,,
China saucer	,, 1	—
Copper wire, S.W.G.22	oz. 2	Special supply.
,, foil	,, 2	,, ,,
Droppers, eye, glass	No. 3	Dispensary.
Enamel cups, ½ pint	,, 4	Special supply.
Funnel	,, 1	Dispensary.
Filter papers, various	,, 24	,,
Forceps, dissecting	pr. 1	,,
Flannel (fomentation)	yd. ¼	,,
Football bladder	No. 1	Special supply or loan from company.
Glass rods	,, 3	Dispensary.
,, tubing	lb. 1	,,
,, prism	No. 1	Improvize or special supply.
,, ground	pce. 1	Grease-proof paper.
Hydrometer	No. 1	Case, milk test. Stewards' store or urine test stand.
Iron rod	,, 1	—
,, wrought	—	Horse shoe.
,, cast	—	—
Kettle	No. 1	—
Lens	,, 1	Special supply.
Magnet, bar	,, 1	,, ,,
Measure, 10-oz.	,, 1	Dispensary.
,, 20-oz.	,, 1	,,
Mortar and pestle	,, 1	,,
Mirror	,, 1	—
Needles, darning	,, 1	Housewife.
,, sewing	,, 1	,,
Spoon	,, 1	—
Sheet, tin	,, 1	Any empty tin.
Silk	sq.yd. 1	Special supply.
,, thread	yds. 2	Dispensary.
Slides, microscope	No. 12	,,
Spirit lamp	,, 1	,, (stand, urine, test).
Scales, druggist	,, 1	,,

Apparatus—*continued*

Description	Approx. quantity for class of six	Suggested source
Test tubes	No. 48	Dispensary.
Tube containing mercury	,, 1	R.A.M. college supply.
,, containing H_2O	,, 1	,, ,, ,,
Tins, empty	—	Kitchen or steward's store.
Thermometers, 0°–110° C.	,, 2	Special supply.
Thermos flask, 01	,, 1	,, ,,
Tin-lids	,, 2	—
Vulcanite rod, 9″ × ½″	,, 1	Special supply.
Whitewash	oz. 1	—
Wire, electric, flex	yd. 1	Special supply.
Weight, ¼ lb.	No. 1	Dispensary.
Watch-spring, old	—	Special supply.
Zinc sheet, 1 ft. × 1 ft.	No. 1	,, ,,

Chemicals

Acid, acetic	oz. 4	Dispensary.
,, boric	,, 4	,,
,, carbolic crystals	,, 2	,,
,, hydrochloric, pure	,, 8	,,
,, nitric, pure	,, 4	,,
,, salicylic	,, 2	,,
,, sulphuric, pure	,, 4	,,
Alum	,, 4	,,
Ammonium carbonate	,, 4	,,
,, chloride	,, 4	,,
Calcium chloride	,, 8	,,
,, carbonate ppt.	,, 8	,,
,, oxide	,, 8	,,
,, sulphate, anhydrous	,, 8	,,
Charcoal	,, 4	,,
Chloroform	,, 6	,,
Chloride of lime	,, 8	,,
Copper sulphate crystals	,, 4	,,
Ferrous sulphate crystals	,, 4	,,
Formaldehyde, 40 per cent.	,, 8	,,
Glycerine	,, 8	,,
Glucose	,, 4	,,
Gum Arabic	,, 4	,,
Ice	,, 8	Steward's store.
Ink	—	Office.
Iron filings (fine)	,, 8	Special supply.
Lead acetate	,, 2	Dispensary.
Liquor, hydrogen peroxide	,, 8	,,
Litmus paper (red and blue)	bks. 4	,,
Liquid paraffin	oz. 4	,,
Liquor ammonium fort.	,, 8	,,

Chemicals—*continued*

Description	Approx. quantity for class of six	Suggested source
Magnesium wire	oz. 2	Special supply.
,, carbonate, Levis	,, 4	Dispensary.
,, ,, Pond	,, 8	,,
,, sulphate	,, 8	,,
Methylated spirit	pt. 1	,,
Mercury	oz. 8	Special supply.
Mercurous chloride	,, 2	Dispensary.
Mercuric chloride	,, 2	,,
,, oxide	,, 2	,,
Moth balls, or Naphthalene	,, 4	,,
Olive oil	,, 8	,,
Pyrogallic acid	,, 4	Special supply.
Potass. chlorate	,, 8	Dispensary.
,, nitrate	,, 4	,,
,, permanganate	,, 8	,,
,, carbonate	,, 8	,,
,, bicarbonate	,, 8	,,
Quinine sulphate	,, 2	,,
,, bisulphate	,, 2	,,
Soap	—	—
Sodium carbonate crystals	,, 8	,,
,, chloride	,, 8	,,
,, bicarbonate	,, 4	,,
,, sulphate	,, 8	,,
,, hydroxide	,, 8	,,
Silver nitrate	,, 1	,,
Starch	,, 2	,,
Sand	—	—
Sugar	oz. 8	Dispensary.
Sealing-wax	sticks 6	Office.
Sulphur flowers	oz. 4	Dispensary.
Sausage skin (or meat membrane)	No. 1	Butcher's shop or steward's store.
Thick oil	oz. 1	Any motor-car oil.
Tinct., balsam	,, 1	Dispensary.
,, asafœtida	,, 1	,,
,, iodine	,, 4	,,
Turmeric papers	No. 12	District laboratory or R.A.M. college.
Turpentine	oz. 4	Dispensary.
Tragacanth	,, 2	,,
Vinegar	,, 4	,,
Zinc, oxide	,, 4	,,
,, carbonate	,, 4	,,
,, chloride	,, 2	,,
,, sulphate	,, 2	,,

APPENDIX 2

ATOMIC WEIGHTS, VALENCIES, ETC.

Name	Symbol	Atomic Weight	Valency	Remarks
Hydrogen	H	1·008	1	Gaseous metal
Helium	He	4·0	0	Rare gas
Lithium	Li	6·94	1	Alkali metal
Beryllium	Be	9·02	2	Alkaline earth, metal
Boron	B	10·83	3	Non-metal
Carbon	C	12·0	4	Non-metal
Nitrogen	N	14·0	1, 2, 3, 4, 5	Non-metal
Oxygen	O	16·0	2	Non-metal
Fluorine	F	19·0	1	Non-metal
Neon	Ne	20·18	0	Rare gas
Sodium	Na	23·0	1	Alkali metal
Magnesium	Mg	24·3	2	Alkaline earth
Aluminium	Al	26·97	3	Metal
Silicon	Si	28·0	4	Non-metal
Phosphorus	P	30·98	3, 5	Non-metal
Sulphur	S	32·06	2, 4, 6	Non-metal
Chlorine	Cl	35·46	1	Non-metal
Argon	A	39·94	0	Rare gas
Potassium	K	39·10	1	Alkali metal
Calcium	Ca	40·09	2	Alkaline earth, metal
Scandium	Sc	45·1	3	Metal
Titanium	Ti	47·9	4	Metal
Vanadium	V	50·95	1, 2, 3, 4, 5,	Non-metal
Chromium	Cr	52·04	2 & 3	Metal
Manganese	Mn	54·95	2 & 3	Metal
Iron	Fe	55·84	2 & 3	Metal
Cobalt	Co	58·95	2 & 3	Metal
Nickel	Ni	58·69	2 & 3	Metal
Copper	Cu	63·55	1 & 2	Metal
Zinc	Zn	65·78	2	Metal
Gallium	Ga	69·72	3	Metal
Germanium	Ge	69·72	4	Metal
Arsenic	As	74·93	3 & 5	Non-metal
Selenium	Se	79·2	2, 4, & 6	Non-metal
Bromine	Br	79·91	1	Non-metal
Krypton	Kr	82·90	0	Rare gas
Rubidium	Rb	85·4	1	Alkali metal
Strontium	Sr	87·6	2	Alkali earth, metal
Yttrium	Ye	88·9	3	Metal
Zirconium	Zr	91·2	4	Metal
Niobium	Nb	93·3	3 & 5	Metal
Molybdenum	Mo	96·0	3, 4, & 6	Metal
Ruthenium	Ru	101·6	2, 3, & 4	Metal
Rhodium	Rh	102·9	2, 3, & 4	Metal
Palladium	Pd	106·7	2 & 4	Metal
Silver	Ag	107·88	1	Metal
Cadmium	Cd	112·4	2	Metal
Indium	In	114·8	3	Metal

Atomic Weights, Valencies, etc.—*continued*

Name	Symbol	Atomic Weight	Valency	Remarks
Tin	Sn	118·7	2 & 4	Metal
Antimony	Sb	121·76	3 & 5	Metal
Tellurium	Te	127·6	2, 4, & 6	Non-metal
Iodine	I	126·93	1	Non-metal
Xenon	Xe	130·2	0	Rare gas
Cæsium	Cs	132·8	1	Alkali metal
Barium	Ba	137·3	2	Alkaline earth, metal
Lanthanum	La	138·9	3	Metal
Cerium	Ce	140·2	3 & 4	Metal
Praseodymium	Pr	140·9	3	Metal
Neodymium	Nd	144·2	3	Metal
Samarium	Sm	150·4	3	Metal
Europium	Eu	152·0	3	Metal
Gadolinium	Gd	157·0	3	Metal
Terbium	Tb	159·2	3	Metal
Dysprosium	Dy	162·4	3	Metal
Holmium	Ho	163·5	3	Metal
Erbium	Er	167·6	3	Metal
Thulium	Tm	169·4	3	Metal
Ytterbium	Yb	173·0	3	Metal
Lutecium	Lu	175·0	?	?
Hafnium	Hf	178·6	?	?
Tantalum	Ta	181·3	4 & 5	Metal
Tungsten	W	184·1	2, 4, 5, & 6	Metal
Osmium	Os	191·0	2, 4, & 6	Metal
Iridium	Ir	193·0	2, 3, & 4	Metal
Platinum	Pt	195·2	2 & 4	Metal
Gold	Au	197·2	1 & 3	Metal
Mercury	Hg	200·6	1 & 2	Metal (liquid)
Thallium	Tl	204·0	1 & 2	Metal
Lead	Pb	207·2	2 & 4	Metal
Bismuth	Bi	209·0	2, 3, 4, & 5	Metal
Niton	Nt	222·0	0	Rare gas
Radium	Ra	225·9	2	Metal
Thorium	Th	232·12	4	Metal
Uranium	U	238·1	4 & 6	Metal

APPENDIX 3

TABLE OF SOLUBILITIES

	Bicarbonate	Borate	Bromate	Bromide	Carbonate	Chlorate	Chloride	Chromate	Cyanide	Ferricyanide	Ferrocyanide	Fluoride	Hydroxide	Hypochlorite	Hypophosphite	Iodate	Iodide	Nitrate	Nitrite	Oxide	Permanganate	Phosphate	Silicate	Sulphate	Sulphide	Sulphite	Sulphocyanide	Thiosulphate	
Aluminium . .	d	i	s	s	d	vs	vs	i	id	s	s	vi	vi			s	s	vs	s	vi	s	i	i	vs	d	s	s	s	Aluminium
Antimony . .	—			d	—		s				i	s	si				d	—		i				d	i			—	Antimony
Arsenic	—			d	—		d					d	s				sd			i					i		d	—	Arsenic
Barium	i	i	i	vs	vi	vs	vs	i	s	i	i	i	s		vs	vi	s	s	vs	i	s	i	vi	vi	d	i	vs	i	Barium
Bismuth	i	i	i	sd	i	s	sd	i	i	i	id	s	i			i	i	d	s	i	id	i	i	d	i	i	s	d	Bismuth
Cadmium	i	i	s	vs	i	vs	vs	i	i	i	i	i	i			i	vs	vs	s	i	s	i		vs	i	i	s	s	Cadmium
Calcium	s	i	s	vs	i	vs	vs	s	s	s	vs	i	i	s	vs	s	vs	vs	vs	i	s	i	i	i	vs	i	s	vs	Calcium
Chromium . .				si			si		—		i	i	si					s		i		i		vs			vs		Chromium
Cobalt	i	i	s	s	i	s	s	i	i	i	i		i			s	i	vs	s	i	s	i	i	vs	i	i	s	s	Cobalt
Copper	i	i	vs	si	i	vs	si	s	i	i	i	s	i		d	i	i	vs	s	i	s	i	i	s	i	i	i	i	Copper
Gold				i			vs		vs			i	i							i					i		—		Gold
Hydrogen (= acids)	—	s	s	s	s	s	s	s	s	s	s	s	s	s	s	s	s	s	s	s		s	is	s	s	s	—	—	Hydrogen (= acids)
Iron	i	i	s	s	si	s	s	s	—	i	i	s	i		s	i	s	s	s	i	s	i	i	s	i	s	s	vs	Iron
Lead	i	i	i	s	i	s	s	i	i	s	i	i	i	d		i	s	s	s	i	i	i	i	i	i	i	id	i	Lead
Lithium	s	s	s	vs	i	vs	vs	vs	s	s	vs	i	i			s	vs	vs	s	s	s	i		vs	s	s	s	s	Lithium
Magnesium . .	s	s	vs	vs	i	vs	vs		—	s	vs	i	i			s	vs	vs	s	s		i		vs	i	s	s	s	Magnesium
Manganese . .	i	i	s	s	i	s	s	i	—	i	i	s	i			i	vs	vs	s	i	d	i		vs	i	i	s	s	Manganese
Mercury (*i.e.* salts)	s	i	s	s	i	s	s	i	s	s	i	s	i	d	d	s	i	d	s	i	s	i		d	i	s	i	d	Mercury (*i.e.* salts)
Nickel				vs	i		s		i	i	i		i					vs		i		i		vs			s	s	Nickel
Platinum				i			si		i				i							i				i			i	—	Platinum
Potassium . .	vs	vs	s	vs	vs	s	vs	vs	s	vs	vs	vs	vs	vs	vs	s	vs	vs	vs	vs	s	vs	s	vs	vs	vs	vs	vs	Potassium
Silver	i	i	i	i	i	s	i	i	i	i	i	s	i	i	i	i	i	vs	s	i	s	i	i	i	i	i	i	i	Silver
Sodium	s	s	s	vs	vs	vs	vs	vs	vs	vs	vs	s	vs	vs	vs	s	vs	vs	vs	vs	s	vs	vs	vs	vs	vs	vs	vs	Sodium
Strontium . .	s	i	vs	vs	i	s	vs	i	vs		vs	s	s			i	vs	vs	s	vs	s	i		i	d	i	s	vs	Strontium
Tin	i	i	s	s	i	s	vs	—	i	i	i	s	i			d	s	d	i	i	d	i		s	i	i	s	s	Tin
Zinc	i	i	s	s	i	s	vs	s	i	i	i	s	i			s	s	s	s	i	i	i	i	s	i	s	s	s	Zinc
Ammonia	vs	s	vs	vs	vs	s	vs	vs	s	s	vs	vs	vs	—	s	s	vs	vs	vs	—	s	vs	s	vs	vs	s	vs	vs	Ammonia

— = compound not known.
vs = soluble above 10 per cent. in water at ordinary temperatures.
s = soluble between 1 and 10 per cent. in water.
i = soluble between 0·1 and 1 per cent. in water.

si = two salts known, one soluble, the other insoluble, *e.g.* under copper, cupric chloride is soluble, cuprous chloride is insoluble.
vi = practically insoluble.
d = decomposed by water.

Where there are several salts, such as antimonious and antimonic hydroxides, two letters are placed to indicate that one is soluble, the other insoluble.

APPENDIX 4

NORMAL AND STANDARD SOLUTIONS

Normal Solutions of Acids

ALL acids contain at least one atom of hydrogen, for example, hydrochloric acid (HCl) contains one, and sulphuric acid (H_2SO_4) contains two atoms.

When an *acid* combines with a *metal* a change occurs, the hydrogen escaping as a gas from the acid and the metal taking its place and forming a salt. For example, hydrochloric acid (HCl) unites with sodium (Na), sodium chloride ($NaCl$) is formed and the atom of hydrogen is set free.

$$Na + HCl = NaCl + H$$

This hydrogen atom is known as *replaceable hydrogen*, because it has been replaced by the sodium atom which occupies the same space in its combination with chlorine as the hydrogen atom did. The weight of the sodium, which will replace one gramme of hydrogen, is termed the equivalent weight of sodium.

Normal Solution

A normal solution of acid contains one gramme of replaceable hydrogen in every litre of the solution.

For example, if to a litre of normal acid solution some zinc is added, bubbles of gas will be given off. This is the hydrogen which is liberated from the acid in the solution, and the chemical action continues until all the hydrogen in the acid is free. If the two substances, namely, zinc and the acid solution, are carefully weighed before they are added to one another, and are again weighed when the gas ceases to escape, then a loss of weight equal to one gramme is observed, and this loss is the weight of the hydrogen replaced by the zinc. We have therefore proof that the solution of the acid was a normal one, for its volume was one litre and the weight of replaceable hydrogen was one gramme.

A decinormal solution contains one-tenth of a gramme of replaceable hydrogen in every litre of the solution, and therefore it is one-tenth the strength of a normal solution.

A centinormal solution contains one hundredth part of a gramme of replaceable hydrogen in one litre, and is one hundred times weaker than a normal solution.

Preparation of a Normal Solution

To make a normal acid solution it is necessary to take the equivalent weight of the acid and dilute it with distilled water up to a litre. These weights have all been determined very accurately by experiment, so that in practice we only need refer to a book on chemistry to obtain them.

The equivalent weights of the different acids, however, can be determined in the following manner.

One hundred grammes of the acid concerned, and some zinc are weighed separately. The zinc is then dropped into the acid, and when bubbles of hydrogen cease to come off from the remaining zinc, the weight is again taken. The difference between the first and second weights is the weight of the hydrogen freed from 100 grammes of the acid, and all that is necessary is to calculate from this figure what weight of acid would liberate one gramme of hydrogen. This weight is the equivalent weight of the acid concerned.

Take for example the preparation of a normal solution of sulphuric acid. The procedure is carried out as described and it is found that the weight of hydrogen replaced is 2·04 grammes. The 2·04 grammes of hydrogen are liberated from 100 grammes of the acid; one gramme of hydrogen would, therefore, be liberated from 49 grammes of the acid.

2·04 grammes of H_2 are set free from 100 grammes H_2SO_4

$$\therefore\ 1 \text{ gramme } H_2 \text{ will be set free from } \frac{100}{2{\cdot}04} \text{ or } 49 \text{ grammes } H_2SO_4$$

Therefore, to get a solution containing 1 gramme of replaceable hydrogen, 49 grammes of pure sulphuric acid must be added to water and diluted up to 1 litre, and this gives a normal solution of sulphuric acid.

A normal solution of hydrochloric acid would be obtained in a similar manner by taking 36·5 grammes of the pure acid, and diluting to 1 litre of water.

These weights are equivalent to 1 gramme of replaceable hydrogen and are known as the *equivalent weights*; for example, 49 grammes of sulphuric acid and 36·5 grammes of hydrochloric each gave 1 gramme of hydrogen.

Normal Solutions of Alkalis

Alkalis may contain hydrogen in their molecule, but it is not replaceable, therefore the method of preparing normal solutions from them cannot be the same as acids.

The method depends on the property *acids* and *alkalis* possess of altering the colour of certain dyes. Litmus is such a dye, and is known as an indicator, because it indicates whether a solution is acid, alkaline, or neutral.

If a solution is acid, litmus becomes red, if alkaline it becomes blue, and if the solution is neither acid nor alkaline, that is *neutral,* then the colour is half-way between red and blue, and appears purple.

Caustic soda (NaOH) is an alkali; when it is added to an acid it destroys or *neutralizes* the acid. So long as the merest trace of acid is present litmus will remain red. There is a point reached, however, when all the acid is neutralized and the litmus changes to purple. This point is known as the *neutral* or *end point*; the amount of alkali just balances the acid present, and any further addition of caustic soda will immediately alter the colour of the litmus to blue, that is, the solution will become alkaline.

How is the amount of alkali required to make a normal solution determined? Again, take an example, caustic soda. A 10 per cent. solution is prepared, that is 10 grammes of caustic soda are dissolved in water and diluted to 100 c.c. This solution is slowly run into a flask containing a litre of a normal acid solution with some litmus in it, until the colour changes from red to purple, which indicates the neutral or end point, when the amount of caustic soda used is noted. Supposing it was 400 c.cs., this contains 40 grammes of caustic soda, because the solution is a 10 per cent. one.

Therefore 40 grammes of caustic soda neutralized or balanced the normal acid solution. Now equal quantities of normal acid and normal alkaline solutions balance one another; and so, if 40 grammes of caustic soda dissolved in a litre of water will balance a similar quantity of a normal acid solution, the caustic soda solution is, therefore, a normal one, and 40 grammes is therefore the equivalent weight of caustic soda, because it is equivalent to the weight of the acid in a normal acid solution, and therefore to 1 gramme of hydrogen.

To repeat: *A normal solution is one which contains the equivalent weight of a substance in a litre of water.*

The equivalent weight of a substance is its weight in grammes, and is equivalent to 1 *gramme of hydrogen.*

Standard Solutions

A standard solution is one made up with a known weight of a substance in order to determine the strength of another substance in solution.

For example, a standard solution of soap is used to determine the amount of hardness in water, and a standard solution of silver nitrate to determine the amount of salt in water.

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APPENDIX 5

TABLE OF INCOMPATIBILITIES

Columns (in vertical headings, same order as the rows): Acacia, Gum. · Acetanilide. · Acids, and Sodium Salts. · Acid Acetylsalicylic. · Carbolic. · Citric. · Gallic. · Hydrochloric. · Hydrocyanic (Prussic). · Nitric. · Nitro-hydrochloric. · Phosphoric. · Salicylic. · Sulphuric. · Sulphurous. · Tannic. · Tartaric. · Aconitine Salts. · Adeps Benzoatus. · Alcohol (Tinctures). · Alum. · Ammon. Acetate. · Benzoate. · Bromide. · Carbonate. · Chloride. · Hydrate. · Amyl Nitrite. · Antimony Salts. · Apomorphine. See Morphine. · Argenti (Silver) Salts. · Arsenic Salts. · Atropine Salts. · Benzamine Lactate. · Bismuth Salts. · Borax. · Butyl Chloral Hydrate. · Caffeine Citrate. · Calcium Salts. · Catechu. · Chloral Hydrate. · Cocaine Salts. · Codeine Salts. · Creta Prep. · Copper Salts. · Diamorphine (Heroin). · Eualgin. · Iron Salts. · Formalin. · Galla. · Hæmatoxylin. · Hexamine. · Homatropine. · Hydrargyrum Salts. · Hyoscyamine Salts. · Iodoform. · Iodine. · Kino. · Krameria. · Liq. Ars. et Hyd. Iod. · Lithium Citrate. · Luminal Sodium. · Magnesium Salts. · Morphine and Apomorphine Salts. · Pepsin. · Phenazone (Antipyrine). · Piperazine. · Plumbi Acetas. · Potassium Salts. · Potassium Permanganate. · Quinine Salts. · Resorcin. · Sarsaparilla. · Sodium Amino-Arsenate. · Spir. Æth. Nit. · Strychnine Salts. · Syr. Ferri Iod. · Syr. Glycerophos. Co. · Tolu (as for Benzoin). · Zinc Salts.

1	Acacia, Gum	
2	Acetanilide	
3	Acids and Sodium salts	
4	Acid, Acetylsalicylic	
5	Carbolic	
6	Citric	
7	Gallic	
8	Hydrochloric	
9	Hydrocyanic (Prussic)	
10	Nitric	All organic materials
11	Nitro-Hydrochloric	
12	Phosphoric	
13	Salicylic	
14	Sulphuric	
15	Sulphurous	
16	Tannic	Albumin, Gelatine
17	Tartaric	
18	Aconitine Salts	
19	Adeps Benzoatus	
20	Alcohol (Tinctures)	
21	Alum	
22	Ammon. Acetate	Alkalies and Carbonates
23	Benzoate	ditto
24	Bromide	ditto
25	Carbonate	ditto and Acids
26	Chloride	ditto
27	Hydrate	ditto and Acids
28	Amyl Nitrite	Alkalies
29	Antimony Salts	
30	Apomorphine. See Morphine	
31	Argenti (Silver) Salts	Organic materials
32	Arsenic Salts	Astringents
33	Atropine Salts	
34	Benzamine Lactate	
35	Bismuth Salts	Carbonates; All acids
36	Borax	
37	Butyl Chloral Hydrate	Alkalies
38	Caffeine Citrate	
39	Calcium Salts	Alkalies
40	Catechu	Gelatine and Alkalies
41	Chloral Hydrate	Alkalies
42	Cocaine Salts	Alkalies
43	Codeine Salts	Alkalies
44	Creta Prep.	
45	Copper Salts	Alkalies
46	Diamorphine (Heroin)	
47	Eualgin	
48	Iron Salts	
49	Formalin	Gelatine
50	Gallae	
51	Hæmatoxylin	
52	Hexamine	Salol; Alkalies
53	Homatropine	
54	Hydrargyrum Salts	Alkalies
55	Hyoscyamine Salts	
56	Iodoform	
57	Iodine	
58	Kino	Gelatine
59	Krameria	Gelatine
60	Liq. Ars. et Hyd. Iod.	
61	Lithium Citrate	Carbonates
62	Luminal Sodium	
63	Magnesium Salts	Acids destroy the Carbonate, Alkalies precipitate, the Sulphate precipitates with strong solutions of Bromides
64	Morphine and Apomorphine Salts	
65	Pepsin	Alkalies; 2
66	Phenazone (Antipyrine)	B. Naphthol; Oil of Wintergreen
67	Piperazine	
68	Plumbi Acetas	Carbonates
69	Potassium Salts	
70	Potassium Permanganate	All organic and reducing substances
71	Quinine Salts	
72	Resorcin	Alkalies
73	Sarsaparilla	Alkalies
74	Sodium Amino-Arsenate	
75	Spir. Æth. Nit.	
76	Strychnine Salts	
77	Syr. Ferri Iod.	
78	Syr. Glycerophos. Co.	Carbonates
79	Tolu (as for Benzoin)	
80	Zinc	Alkalies

A cross indicates that the substance in the horizontal line is incompatible with the substance in the vertical line.
Where the symbol of a substance replaces the cross, this indicates that that particular compound is incompatible, while other compounds of the substance are compatible.
2 incompatible with strong Hydrochloric, compatible with 2 per cent.

In [illegible]