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PHYSICS AND CHEMISTRY FOR NURSES

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By AMY ELIZABETH POPE

ESSENTIALS OF DIETETICS

A QUIZ BOOK FOR NURSES

ANATOMY AND PHYSIOLOGY FOR NURSES

MEDICAL DICTIONARY FOR NURSES

(WITH ANNA CAROLINE MAXWELL)

PRACTICAL NURSING

PHYSICS AND CHEMISTRY

Physics and Chemistry for Nurses

By

Amy Elizabeth Pope

Graduate of the School of Nursing of the Presbyterian Hospital,
in the City of New York; Special Diploma in Education
from Teachers College, Columbia University, New
York; Formerly Instructor in the School of Nursing, Presbyterian Hospital; Instructor in the
School of Nursing, St. Luke's Hospital,
San Francisco, Cal.

Author of "Quiz Book of Nursing," "Anatomy and Physiology
for Nurses," "A Medical Dictionary for Nurses," and,
with Anna Maxwell, of "Practical Nursing"

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PREFACE

INSTRUCTION in hospital-housekeeping is gradually attaining a well-deserved place in the curriculum of schools of nursing. Since for nurses to have a proper understanding of methods of cleaning, laundry work, cooking and diet is of infinite value both to the hospital and to the nurses, especially those nurses who wish to prepare for executive positions in hospitals and for social-service work.

To provide a foundation for such knowledge, by explaining the principles of the chemical and physical properties upon which the various actions that occur in cleaning, cooking, digestion, metabolism, etc., depend, was one reason for compiling this book. A second purpose was to explain important chemical and physical processes constantly referred to in physiology, materia medica, and the other studies included in the school of nursing curriculum.

Also, the book contains many important applications of the chemical and physical processes described to disinfection, cleaning, cooking, and other procedures of interest to nurses.

In compiling this work, I have consulted all available books of chemistry and physics published within recent years, but those which were most frequently referred to are mentioned in the bibliography at the end of this volume, one or more of which books on each subject should be in every school-of-nursing reference library. The majority of experiments here described, I have seen performed in the Household Chemistry Class at Teachers College, Columbia University, New York, or in the Domestic Science Classes at the State Normal College, Santa Barbara, California.

There are probably more experiments given than there will be time for pupils in the majority of schools to perform, but the book has been compiled with the intent that any or all of the experiments can be omitted without interfering with the value of the lessons. In fact, many of the experiments described were inserted to give the students a knowledge of how the facts they demonstrate were ascertained and to illustrate facts referred to, rather than that the experiments should be performed, except by those who are especially interested.

The tests for food adulteration and the illustrations of calorimeters were taken from the different U. S. Department of Agriculture publications mentioned in the text, with the kind permission of the Secretary of Agriculture.

AMY E. POPE.

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Physics and Chemistry for Nurses

CHAPTER I

DIRECTIONS FOR LABORATORY WORK

Names and Nature of Utensils and Chemicals Used for the Experiments Given in this Book—Necessary Care of Utensils—Laboratory Maxims—Laboratory Methods—Weights and Measures

LABORATORY work, properly performed, is of the greatest assistance in gaining an understanding of chemical and physical action, but when carelessly done, it is a waste of time and material and is often dangerous. To get the best results, the student must, from the beginning of her class work, be familiar with the nature and the use of her apparatus and chemicals and know how to work with and care for all her utensils; therefore, this first chapter should be read with attention.

Names and Nature of the Principal Utensils Required for the Experiments Given in this Book

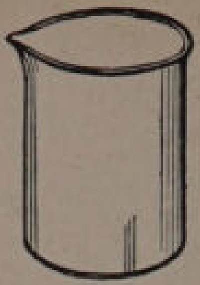
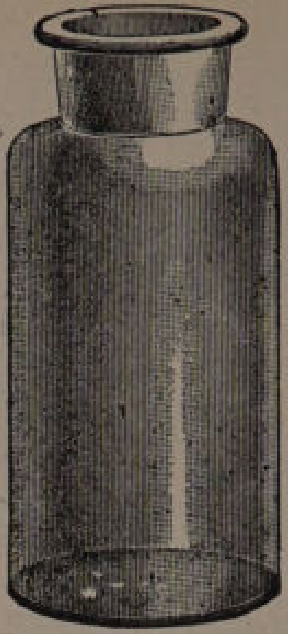


Fig. 1. Beaker.

Beakers are made of thin highly annealed glass and can be used for boiling liquids in if wire netting is placed between them and the flame.



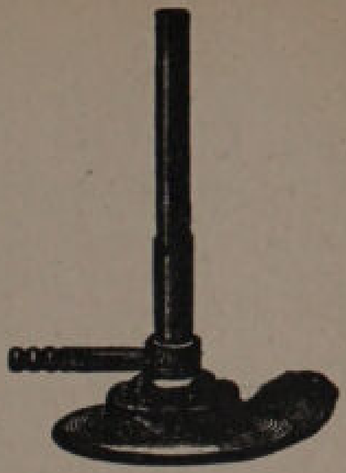


Fig. 3. Bunsen Burner.

Fig. 2. Bottle.

Used to hold cold substances and to collect gas. As it is of thick glass, it cannot be used for boiling liquids.

Fig. 4. Crucible.

This is made of fused silica. It will stand a very high temperature even dry solid substances can be heated in it.

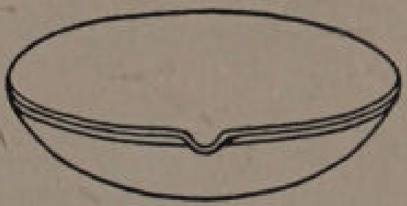


Fig. 5. Evaporating Dish.

Liquids can be heated in porcelain evaporating dishes, if wire gauze is placed between the dishes and the flame.

Fig. 6. Fermentation Tube.

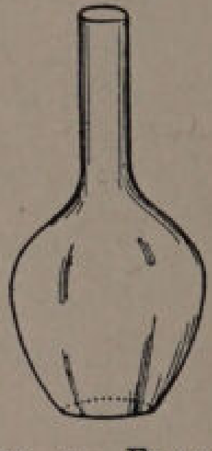


Fig. 7. Flask.

Flasks being of thin glass, liquid can be boiled in them if wire netting is put between them and the flame.

Fig. 8. Erlenmeyer Flask.

This is made of the same kind of glass as the flask shown in Fig. 7.

Fig. 9. Funnel.

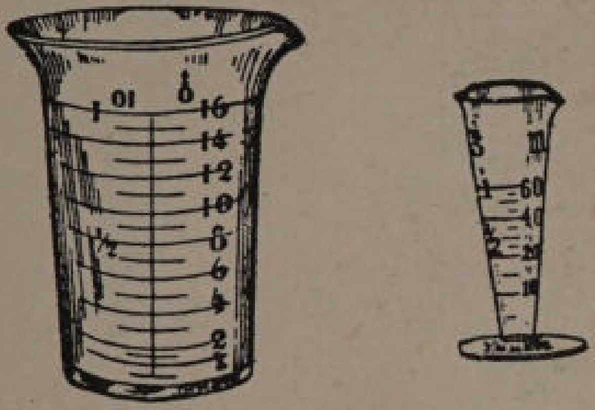


Fig. 10. Graduated Measures.

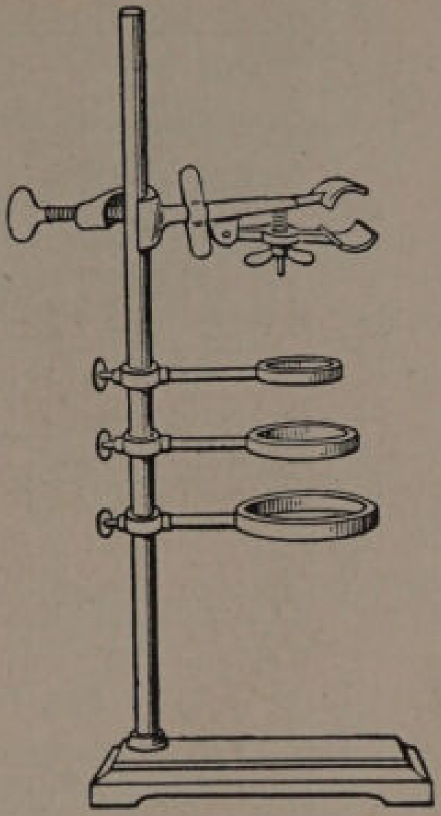


Fig. 11. Iron Stand.

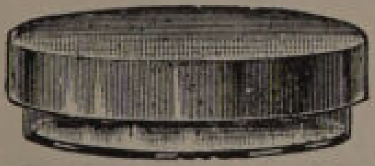


Fig. 12. Petri Dish.

Fig. 13. Pipet.

A pipet made as described on page 15 will answer the purpose for class work.

Fig. 14. Platinum Wire in Glass Rod.

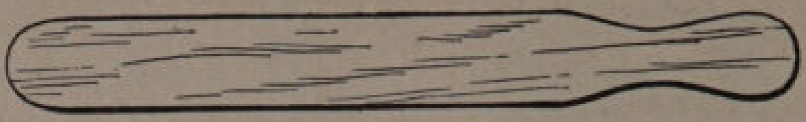


FIG. 15. SPATULA (*Horn*).

Metals being injured by many chemicals, horn spatulas are used for mixing dry substances, and glass rods for stirring liquids instead of spoons.

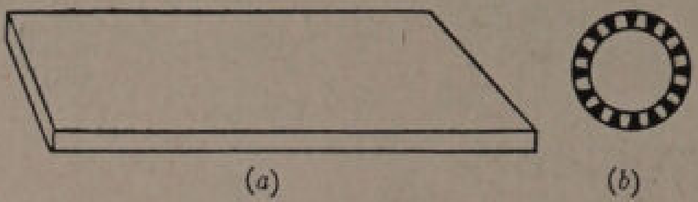


FIG. 16.

(*a*) slide, (*b*) cover glass for use with microscope. For description of microscope, page 98.

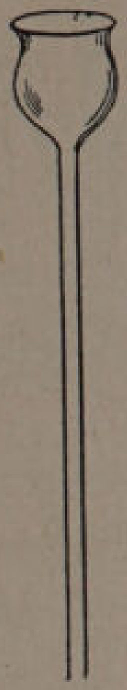


FIG. 18. THISTLE TUBE.

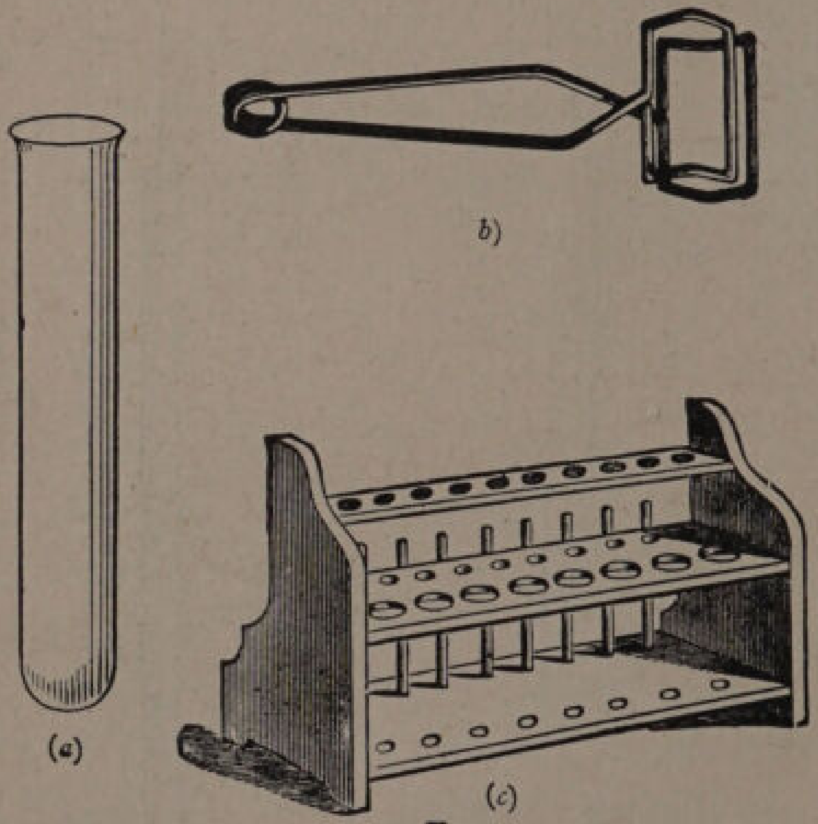


FIG. 17.

(*a*) Test tube.
(*b*) Test-tube holder.
(*c*) Test-tube rack.

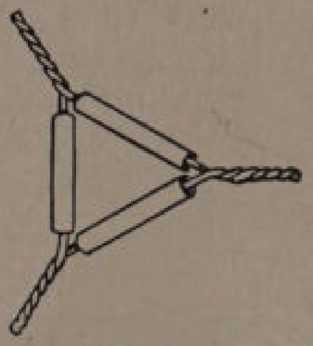


FIG. 19. TRIANGLE.

FIG. 20. WATCH CRYSTAL.

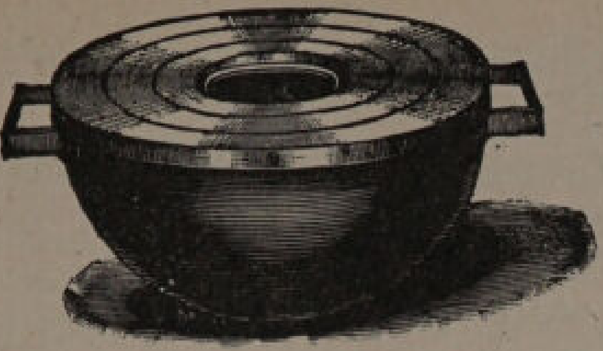


FIG. 21. WATER BATH.

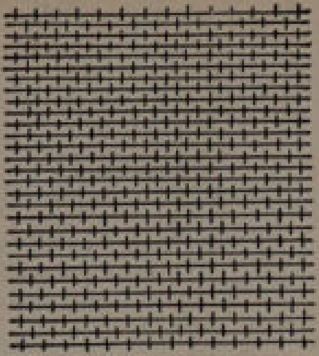


FIG. 22. WIRE GAUZE.

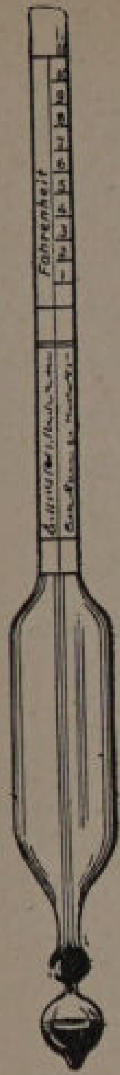


FIG. 23. HYDROMETER.

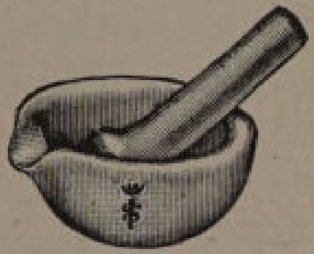


FIG. 24. MORTAR AND PESTLE.

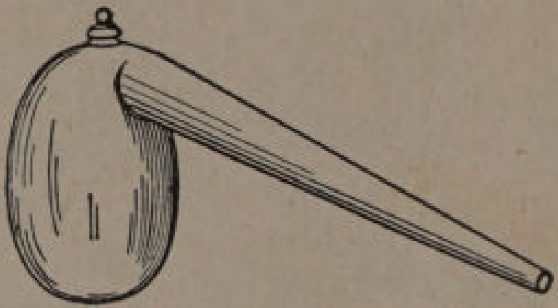
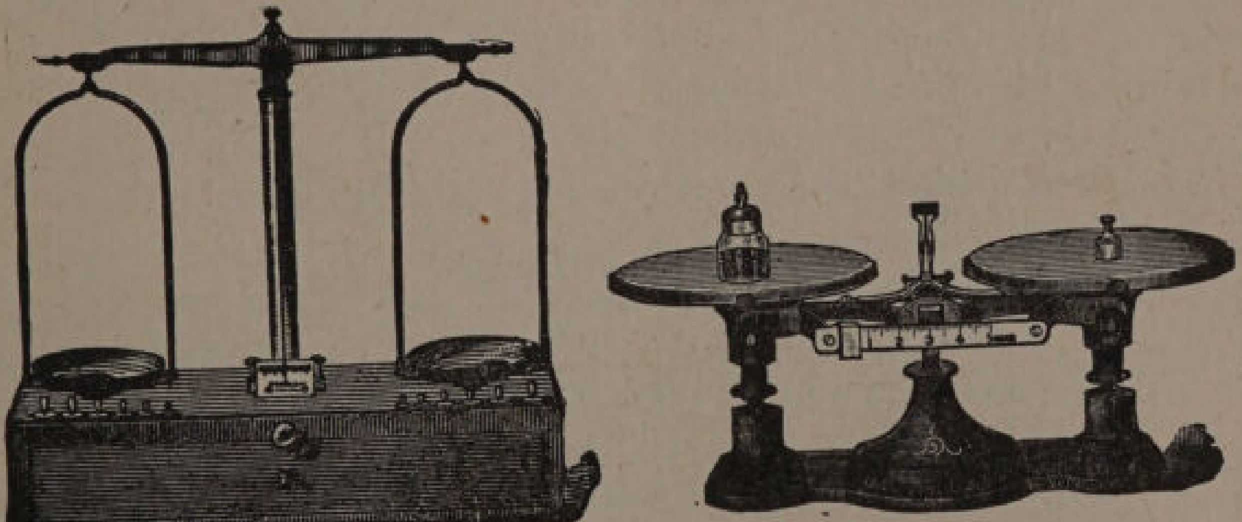


FIG. 25. RETORT.



(a) (b)

FIG. 26. SCALES.

(a) Balance used for weighing small quantities of light-weight material. Trip scales, as (b), are not highly enough sensitized for this purpose; they are, however, necessary for heavier material.

(Figs. 23–26 are mentioned last as they will not be found in the individual equipment, there being, usually, only a limited number of these articles for general use.)

In addition to the articles illustrated, there will be needed for each individual equipment:

A small asbestos mat.

Corks, preferably of rubber, to fit a test tube, a flask, and a heavy glass bottle. The test-tube cork should have one hole through the center. There should be two corks for the flask and bottle, one with one hole and the other with two holes through it. These holes must be exactly the size of the glass tubing that is used.

Glass and rubber tubing about ¼ inch in diameter.

A piece of glass about ⅛ inch thick and 3 inches square.

A tin or enamel dish about 2½ inches deep and 10½ inches in diameter.

Forceps.

Scissors.

File for cutting glass.

Glass rod. A very good substitute for the solid rods usually bought can be made as described on p. 14.

Filter paper.

Litmus paper, blue and pink.

Thermometer, chemical.

A piece of copper wire about 9 inches long.[1]

Bunsen Burner

This burner was so called after the inventor, a German chemist, named *Bunsen*. The principles of its arrangement have been adapted to nearly all forms of gas-stove burners.

Experiment 1. Object: To study the Bunsen burner.

Procedure: Take the burner to pieces and examine

---

[1] This list does not include the apparatus for milk testing, see page 385, nor the articles required for urinalysis.

its different parts. Put the parts together again and connect the burner with the gas pipe. Turn on and light the gas; to do so, turn on a full current of gas and hold the lighted match about two inches above the top of the burner, then regulate the flow of gas so that there will be no waste. Close and open the air holes and notice the difference in the flame. Extinguish the flame by pinching the tube. Light the gas at the air holes. Extinguish the flame and relight the gas in the usual manner. Arrange the air holes so that there is a yellow flame. Do this by turning the small outer tube. Hold a piece of glass tubing in the flame.

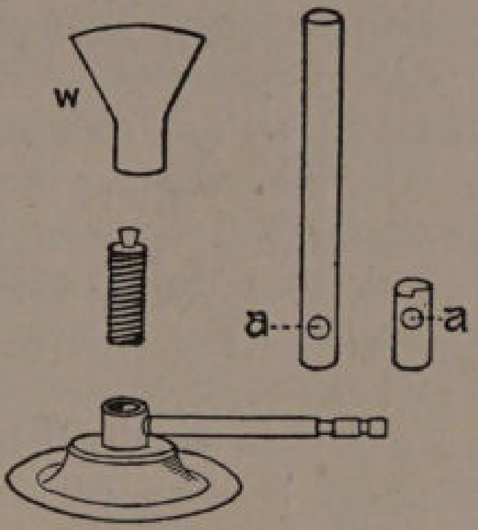


FIG. 27. PARTS OF A BUNSEN BURNER.

(*a*) Air holes.

(*w*) Wing top. This is used when a wide flame is needed.

How long does it take to get the tubing soft enough to bend?

Is the flame luminous, *i. e.*, does it give light?

Is there a deposit formed on the glass tubing?

Is there much or little air entering the air holes of the burner?

Arrange the air holes so that the flame becomes blue.

In what way does this arrangement differ from that required for a yellow flame?

Is there a deposit on the tubing?

How long does it take to get the tubing soft enough to bend?

Is the flame more or less luminous than the yellow flame?

Is there more or less air entering the air holes than there was for the yellow flame?

Which flame will be the best for cooking over?

Mention two objections to the other flame for cooking.

State a reason why the wrong kind of a flame sometimes occurs in a cooking stove.

Why does the gas sometimes *strike back* and burn in the pipe?

The pupils should endeavor to answer the preceding questions before reading the following paragraphs.

When the gas is turned on it flows through the tube, and, as it passes the holes at the bottom of the tube, which are provided to supply air, it will, if the holes are open, draw the air through them. The result is a blue flame, which is hot, but non-luminous. This is often spoken of as the *Bunsen flame*. It is the flame used for laboratory work, unless otherwise specified, and it is also the correct flame for cooking over.

If the air holes are closed, there will not be sufficient oxygen admitted to the tube to unite with all the carbon of the gas, consequently, some of the latter will be separated in the flame by the heat and, at the same time, its temperature will be raised to a white heat and thus it will become luminous. If anything is held over the flame, this unoxidized carbon will be deposited upon it. This is often spoken of as *soot*. The flame will not be as hot as the non-luminous blue flame, because there will not be as much oxidation (the combining of oxygen with a substance) going on, on account of the comparative smallness of the air supply. If too much air gets into the tube, it and the gas will unite to form an explosive mixture and, if this is ignited, the force of the explosion will cause the flame to *strike back* to the opening where the gas enters the tube.

Nature of the Bunsen Flame

Experiment 2. Object, to study the nature of the Bunsen flame.

Procedure: Secure a non-luminous flame and notice that it has three distinct color zones. Hold a piece of bright copper wire in the tip of the outer zone. Notice the discoloration that occurs. Hold the discolored part of the wire near the base of the central zone. Observe that the discoloration acquired in the tip of the outer zone disappears. Why does it do so?

Experiment 3. Take a piece of tubing, bent as in Fig. 29, and hold one end with its opening in the tip of the lower zone. After a few minutes apply the flame of a lighted match to the opening of the free end of the tube. Explain the result.

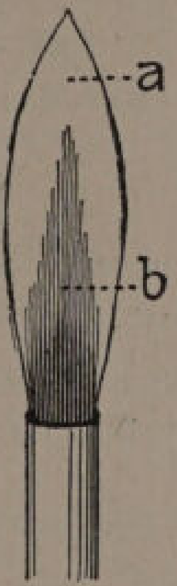


FIG. 28. PARTS OF A FLAME. (*a*) Oxidizing flame. (*b*) Reducing flame.

The dark zone at the bottom of the flame is dark because of the gases issuing from the burner. In the middle zone, due to the heat, these gases are rapidly uniting with the oxygen of the air, but there is still a large quantity of gas present uncombined with oxygen. In the outer zone are the exceedingly hot products of combustion and some uncombined oxygen that is being separated from the air. Due to the presence of this uncombined oxygen, anything that will unite readily with oxygen, as copper does when it is heated, will become oxidized (*i. e.*, it will unite with oxygen) when held in the outer zone of the flame. This zone, therefore, is called the *oxidizing flame*. In the central portion of the flame the hot uncombined gases will remove oxygen from a substance and this part of the flame is therefore called the *reducing flame*, reduction being the word used by chemists to denote the removal of oxygen from a substance. Though there is a larger quantity of uncombined gas in the lower than in the central zone, reduction will not take place there, the heat not being sufficient to allow of the removal of oxygen by the gas.

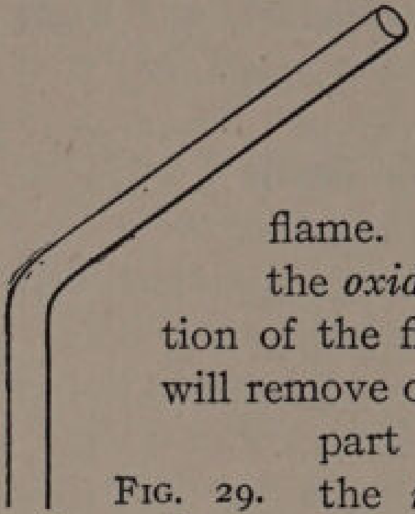


FIG. 29. Shape of tube for Experiment 3.

Precautions Necessary in the Use of Utensils

Do not leave empty vessels of any kind over a flame.

When about to heat anything in glass or porcelain utensils, place the burner in position, light the gas, and regulate the flame before putting the utensil over it.

Use a low flame at first.

Never heat glass or porcelain utensils that are wet in parts not in contact with the contained liquid. The reason why they may break if this is done will be seen in the sections on expansion and contraction.

Always rotate test tubes in the flame until all portions that are to be heated are uniformly hot.

When heating liquids in a test tube, hold the tube slantwise as in Fig. 30.

Put a piece of wire gauze between the flame and glass beakers and flasks and porcelain dishes, except crucibles. The wire spreads the flame and thus prevents any one part of the utensil becoming much more highly heated than another. The reason why this is necessary, like many of the other precautions essential in the care of glassware, will be found in the sections on contraction and expansion. Crucibles can be placed in a triangle directly over the flame, if the heat is applied gradually.

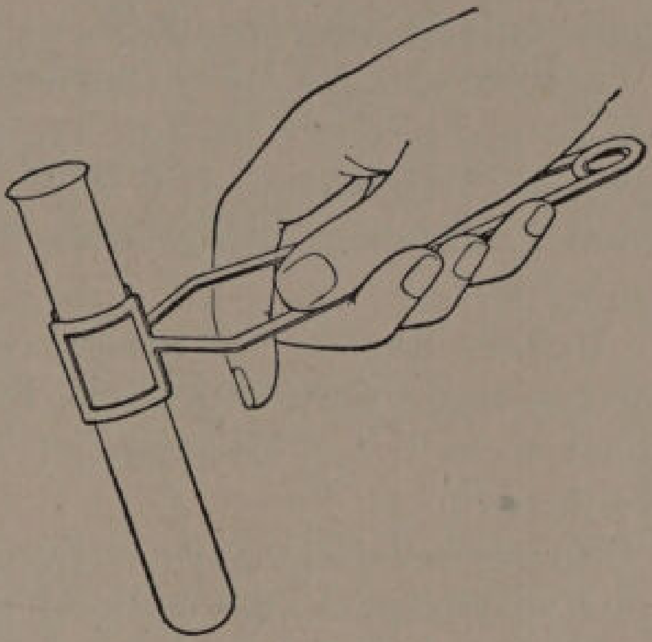


FIG. 30. MANNER OF HOLDING TEST TUBE.

After removing flasks, etc., from the fire, place them on a warm asbestos pad, never on a cold surface.

Clean all utensils thoroughly after use and always be sure that they are clean before use, otherwise, experiments may be spoiled.

Numbering test tubes.—In order to avoid confusion, it is often necessary to number test tubes, etc. This is best done by cutting Dennison's labels into small pieces (about ¼ inch square), numbering these, and sticking them on the utensil.

Points to Remember about Chemicals

They are costly, do not waste them.

Many chemicals are inflammable; do not bring those which are near fire.

Do not inhale vapors arising in the course of experiments or from chemicals.

Remember that many chemicals are very corrosive, so do not spill any and do not put the bottles away wet on the outside. If a strong acid is spilt on the flesh or clothing, apply ammonia or other alkali at once.

Unless otherwise directed, when mixing strong acids—as concentrated nitric, hydrochloric, and sulphuric—with a diluent—*e. g.*, water—add the acid to the diluent.

Remember that many chemicals combine to form explosive mixtures and, therefore, be careful not to form the habit of bending over your work, and obey all precautions given in the instructions.

Laboratory Maxims

Be accurate, be methodical, be careful, be clean, be observing, remember that though the results of some experiments are very pronounced, others are difficult to discern.

When performing experiments, keep your notebook at hand and record, at once, weights and measures and everything of importance observed.

Be on the alert to see, smell, and hear, but do not taste unknown substances unless directed to.

Laboratory Methods

The Preparation of Glass Tubing

To cut glass tubing.—Make a scratch on the tubing where it is to be divided with a triangular file; take the tube in both hands with the thumbs almost together, as in Fig. 31, opposite the scratch, and push slightly upward with the thumbs. If the tube does not break readily, make the scratch deeper, do not use force. When cutting hard glass or heavy tubing, put a handkerchief or piece of cloth between the tubing and the hands, in case the glass should splinter. After glass is severed in this way, the severed ends should be rotated in a flame for a few minutes in order to remove sharp edges.

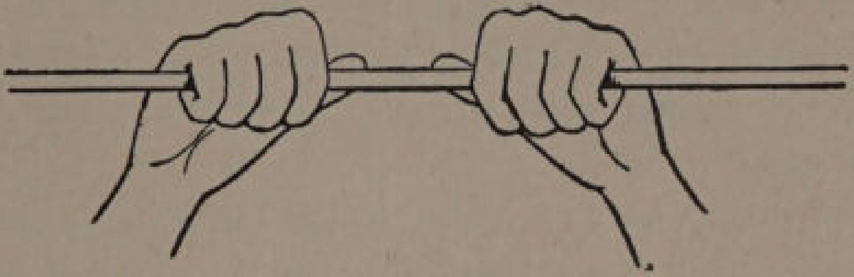


FIG. 31. MANNER OF HOLDING GLASS TUBE WHILE BREAKING IT.

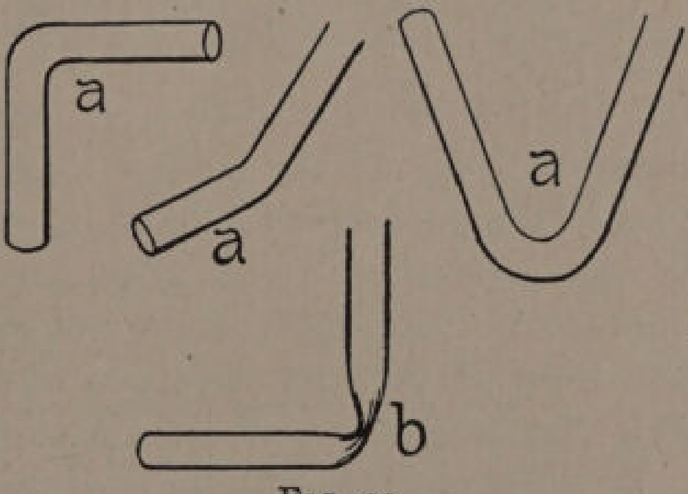


FIG. 32.
(*a*) Good bends. (*b*) A poor bend.

Bending glass tubing.—Necessary precautions: Heat and cool the tubing slowly, therefore, coat the bend with soot by holding it in the yellow flame and do not lay the tubing on a cold surface while it is hot. Rotate the tubing while heating it, especially at first, so as to prevent cracking by uneven heating. Apply the heat continuously and do not attempt to bend the tubing until the heat has rendered it perfectly pliable. Be careful while bending the tube to keep the bend round; if the curve is so sharp that the tubing is flattened at any point the passage of liquid or gas will be interfered with, also such bends are brittle and break easily.

Procedure: Use a wing-top burner. Secure a yellow flame. Keep this flame for thin-walled tubing, but for that with a thick wall, change to a colorless flame as soon as the tubing is slightly coated with soot at the part which is to be bent. Hold the tubing between the thumb and fingers and rotate it in the flame until it is pliable; then remove it from the flame and bend it into the required shape. It is well to keep the tube on an asbestos mat while bending it.

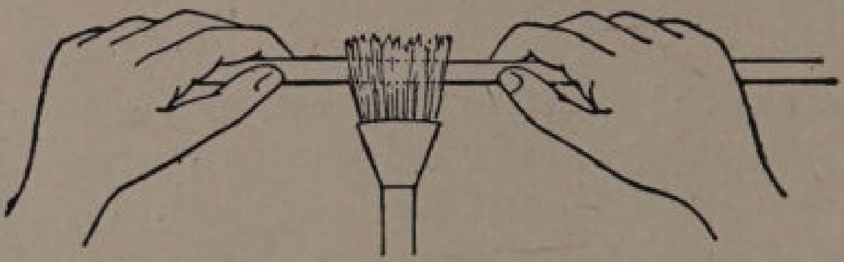


FIG. 33. HOLDING GLASS TUBING IN THE FLAME.

To draw out glass tubing.—Rotate the central portion of the tubing in the flame until it is soft. Then make traction on both ends until the points are of the desired diameter.

Let the glass become perfectly cold and then cut it as previously directed. Tubing prepared in this way can be used for pipets.

To make glass stirring rods.—Heat and draw the tubing as directed in the preceding paragraph, but instead of cooling and cutting the point, draw the glass apart and fuse the ends in the flame until there is no opening. This will result in two tubes each closed at one end. If a bulb is wanted at the end of a rod, rotate the closed end in the flame until it is soft, remove it from the flame, and, at once, blow into the open end. Continue blowing until the bulb is the desired size.

FIG. 34. GLASS TUBE DRAWN TO A POINT, READY TO CUT.

Filtering

Filtering is a process by means of which finely divided solids are removed from liquids. It consists in making liquids pass through a substance that is sufficiently porous not to interfere with the passage of

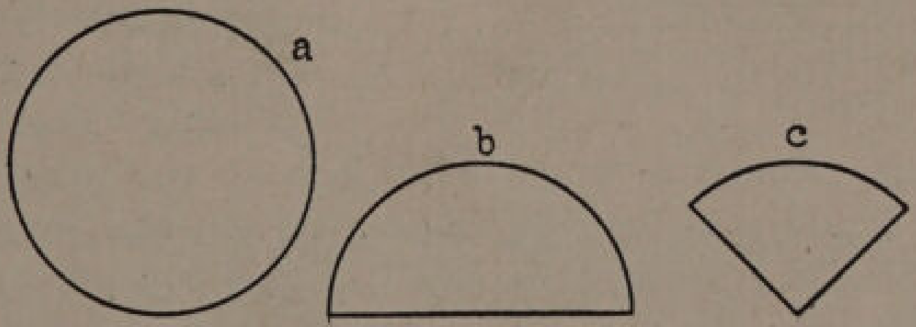


FIG. 35. FOLDING FILTER PAPER TO FIT FUNNEL.
(*a*) Filter Paper, (*b*) first folding, (*c*) second folding.

liquid matter, but dense enough to prevent or impede the passage of solids suspended in the liquid. The methods of filtering will be discussed in Chapter XV. In the laboratory, filtering is usually done with filtering paper. This is folded as shown in Fig. 35, and placed in a funnel. When folding the paper, be careful not to press it at the point or it will tear. Place the paper in the funnel with three thicknesses on one side and one on the other. Do not allow the paper to project above the funnel and fit it into the funnel in such a manner that its entire surface is supported, otherwise the weight of the liquid may tear the paper. When the liquid to be filtered is thick, it is usually better to moisten the paper with distilled water before pouring in any of the liquid. Also, it is well to pour the liquid down the side of a rod held with its point upon the single layer of paper, as in Fig. 36.

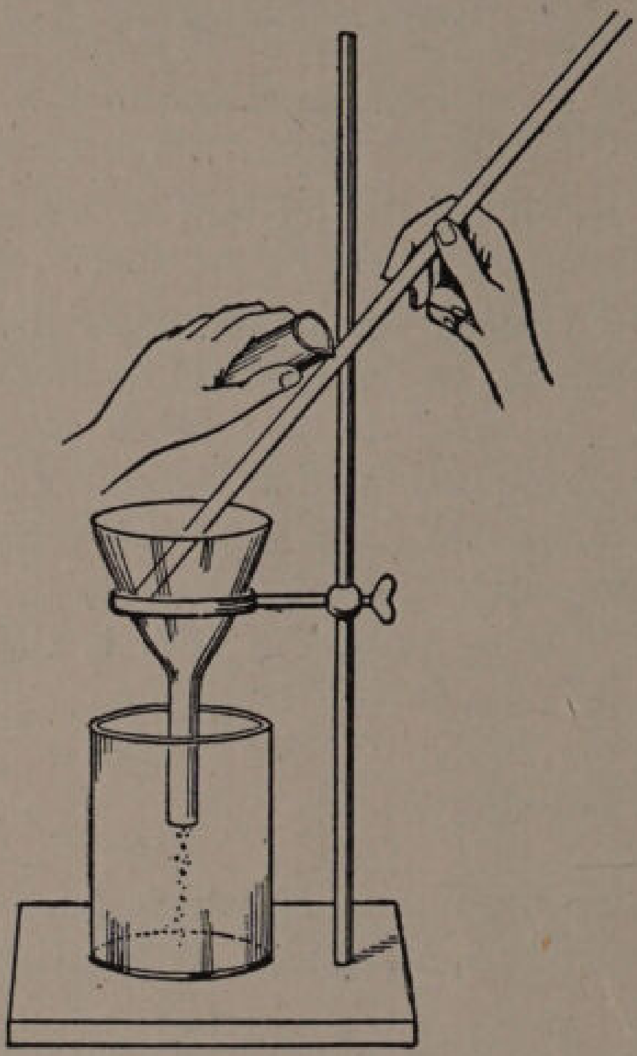


FIG. 36. FILTERING.

To introduce a powdered substance into a narrow necked flask or a test tube.—Fold a strip of paper as shown in Fig. 37, place the powder in one end of the trough thus made; introduce the paper into the tube, as in Fig. 37, turn the paper so that the powder will be deposited where desired; withdraw the paper.

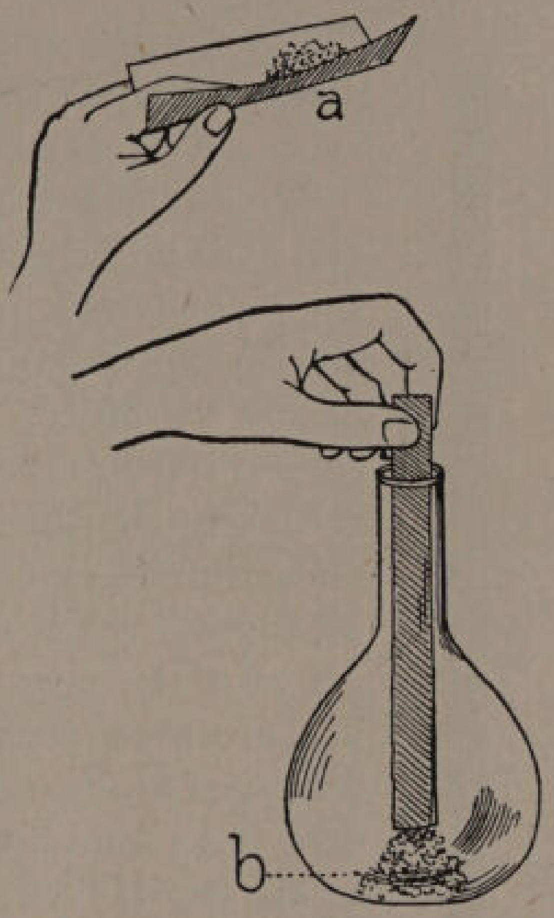


FIG. 37. PUTTING A POWDER INTO A NARROW-NECKED FLASK.

(*a*) Paper trough with powder. (*b*) Powder deposited in flask.

Weighing

When weighing chemicals, put them on a piece of paper or in some utensil, as a watch crystal, beaker, etc.; do not put them on the unprotected pan of the scales for many chemicals will corrode or otherwise injure the pan. The weight of this paper or utensil must be either noted, recorded, and afterward deducted from the total weight of utensil and chemical, or else an object of equal weight must be placed in the other pan. The weights are put in the center of the right-hand pan.

Pupils should form the habit of handling weights with forceps, for when small amounts are being weighed, even moisture from the fingers will interfere with an absolutely correct result.

Measuring

Liquids are measured in graduated glass measures, pipets, and burets.

FIG. 38. BURET

When using measures, hold the measuring glass so that the lowest point of the meniscus[1] is on a level with the eye.

To use a pipet, hold it between the thumb and the middle finger of the right hand; place the pointed end in the liquid and the other end in the mouth and suck up the liquid nearly to the top of the tube, but be very careful not to draw it into the mouth; while removing the tube from the mouth, quickly place the index finger over the top. Then, by slightly releasing the pressure of the finger over the opening, either allow the number of drops required to escape, or, if the pipet is a graduated one, allow the liquid to drop from the tube, back into the bottle, until the top of its column is on a line with the mark of the quantity required. Then hold the pipet over the utensil into which the liquid is to be put and release the finger over the opening; the pressure of the air entering the tube will force the liquid from it.

When using a buret, pour the liquid to be measured into the buret tube, place a receptacle under it, and, carefully, release the pressure of the stop cock on the rubber tubing until the amount of liquid required has passed into the receptacle.

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[1] The upper surface of a fluid column, as that of mercury in a thermometer. The lowest point, *i. e.*, in the center, is considered the correct height of the column of liquid.

List of Chemicals, Solutions, and Reagents Needed for Experiments

Acetic acid. CH_3COOH. Commercial acetic acid is about 36%. Glacial acetic acid is 99%.

Acetic anhydrid. $C_4H_6O_3$.

Alcohol, amyl. $C_5H_{11}OH$.

Alcohol, ethyl. C_2H_5OH.

Alum. $KAl(SO_4)_2$.

Ammonia water. NH_4OH. This as purchased is about 28%. Known also as ammonium hydroxid.

Ammonium molybdate solution. $(NH_4)_2MoO_4$. This is prepared as follows: Add 25 gm. of molybdenum oxid (MoO_3) to 100 cc. of NH_4OH and stir the solution until the powder is dissolved. Cool and add 250 cc. of nitric acid.

Ammonium oxalate. $(NH_4)_2C_2O_4$.

Barium chlorid. $BaCl_2$.

Benzoic acid. $C_7H_6O_2$.

Borax. $Na_2B_4O_7$. $10H_2O$.

Boric acid. H_3BO_3.

Calcium carbonate. $CaCO_3$.

Calcium chlorid or chlorid of lime. $CaCl_2$.

Calcium sulphate. $CaSO_4$.

Carbon disulphid. CS_2. This is very inflammable. It is only required for Experiment 75.

Chloroform. $CHCl_3$.

Copper sulphate. $CuSO_4$.

Fehling's solution. This is prepared as follows:

Solution A: Dissolve 34.64 grams of copper sulphate in 400 cc. of distilled water. When this is dissolved, add enough water to make 500 cc. of solution. Solution B: Dissolve 180 grams of Rochelle salt in 500 cc. of sodium hydroxid solution, 10 %.

Ferric alum. $KFe(SO_4)_2.\ 12H_2O$.

Fibrin. Dried blood fibrin can be obtained from dealers in chemical apparatus.

Hydrochloric acid. Known also as muriatic acid. HCl. The concentrated solution is 40%. For the dilute solution add 1 volume of the concentrated acid to 4 volumes of water.

Iodin. I. Tincture of.

Lead acetate. $Pb(C_2H_3O_2)_2$.

Litmus powder.

Logwood, tincture of.

Mercuric chlorid, corrosive sublimate, or bichlorid of mercury. $HgCl_2$.

Millon's solution. Make as follows: Dissolve 5cc. of mercury in 95 cc. of concentrated nitric acid and, when action ceases, add twice the volume of cold water.

Nitric acid. Known also as aqua fortis. HNO_3. Concentrated nitric acid is 70%. For the dilute solution add 1 volume of the concentrated acid to 4 volumes of water.

Oxalic acid. $(CO_2H)_2$.

Pancreatin solution. This can be made by dissolving 2 grams of the commercial pancreatin in 300 cc. of warm—not hot—water. Or it can be prepared as follows from the fresh pancreas of a pig: Free the pancreas from fat as far as possible; mince it, weigh it, put in into a flask, and add three times its weight of distilled water and its own weight of alcohol. Shake the flask vigorously and then let it stand for three days at room temperature, shaking it occasionally. Then strain the solution first through muslin and afterward through filter paper. Measure the filtrate and, for every liter, add 1 cc. of strong HCl. Let this stand for a week and then filter. Pour the filtrate into a flask that can be tightly corked. It will keep indefinitely, the alcohol acting as a preservative. The solution contains trypsin and amylopsin, but no lipase. See Chapter XXII.

Pepsin solution. This can be made from the commercial pepsin. The percentages required are given in the experiments, Chapter XXII., and the solution should be prepared on the day that it is wanted, as it does not keep.

Potassium ferrocyanid or yellow prussiate of potash. $K_4FeC_6N_6$.

Potassium permanganate. $KMnO_4$.

Potassium thiocyanate or potassium sulphocyanid. KSCn.

Salicylic acid. $C_6H_4OH.COOH$.

Silver nitrate. $AgNO_3$. Solution 2%.

Soap solution. Make as follows: Shave and dissolve 50 grams of white castile soap in 1 liter of hot water. Filter the resulting solution.

Sodium carbonate. Na_2CO_3.

Sodium chlorid. NaCl.

Sodium hydroxid. NaOH.

Starch solution. Blend one teaspoonful of starch with one tablespoonful of cold water and then add slowly, while stirring, 250 cc. of boiling water. Boil the solution for three minutes. This should not be prepared long before it is required for use.

Sulphuric acid. H_2SO_4. Concentrated 94%. For the dilute solution add 1 volume of the commercial acid to 4 volumes of water.

Tumeric paper.

Points to Remember when Making Solutions

When diluting the acids, add the acid to the water and use cold water.

When dissolving solid substances, use boiling water unless otherwise specified. Two important exceptions are pepsin and pancreatin; for these use warm, but not hot, water.

When calculating the amount of a chemical to use remember that 1 per cent. means one part in one hundred. Therefore to make a 1 per cent. solution use 1 gram or 1 cc. of the substance to be dissolved to 100 cc. of water or whatever solvent is to be used. For a 5 per cent. solution, use 5 grams or 5 cc. to 100 cc. For a 10 per cent., use 10 grams or 10 cc. to 100 cc. For a $\frac{1}{10}$ of a 1 per cent. use 1 gram or 1 cc. to 1000 cc. For a $\frac{2}{10}$ of a one per cent. use 2 grams or 2 cc. to 1000 cc. of solvent.

Unless otherwise specified the strengths of the solutions for the experiments are to be as follows: Dilute acids, diluted according to the directions already given; concentrated ammonia; 10% solutions of the other hydroxids and the salts.

METRIC WEIGHTS

Milligram	=	.001	of a gram
Centigram	=	.01	“ “ “
Decigram	=	.1	“ “ “
Gram	=	Principal unit	
Decagram	=	10	grams
Hectogram	=	100	“
Kilogram	=	1000	“
Myriogram	=	10,000	“

COMPARATIVE VALUES OF APOTHECARIES' AND METRIC FLUID MEASURES

Minims	*Cubic Centimeters*	*Minims*	*Cubic Centimeters*	*Fluid Ounces*	*Cubic Centimeters*	*Fluid Ounces*	*Cubic Centimeters*
1	0,06	25	1.54	2	59.20	21	621.00
2	0.12	30	1.90	3	89.00	22	650.00
3	0.18	35	2.16	4	118.40	23	680.00
4	0.24	40	2.50	5	148.00	24	710.00
5	0.30	45	2.80	6	178.00	25	740.00
6	0.36	50	3.08	7	207.00	26	769.00
7	0.42	55	3.40	8	236.00	27	798.50
8	0.50			9	260.00	28	828.00
9	0.55	*Fluidrams*		10	295.70	29	858.00
10	0.60	1	3.75	11	325.25	30	887.25
11	0.68	1¼	4.65	12	355.00	31	917.00
12	0.74	1½	5.60	13	385.00	32 or 1 quart	946.00
13	0.80	1¾	6.51	14	414.00	48	1419.00
14	0.85	2	7.50	15	444.00	56	1655.00
15	0.92	3	11.25	16 or 1 pint	473.11	64	1892.00
16	1.00	4	15.00	17	503.00	72	2128.00
17	1.05	5	18.50	18	532.00	80	2365.00
18	1.12	6	22.50	19	562.00	96	2839.00
19	1.17	7	26.00	20	591.50	112	3312.00
20	1.25					128	3785.00
		Fluidounce					
		1	30.001				

COMPARATIVE VALUES OF APOTHECARIES' AND METRIC WEIGHTS

Grains	*Grams*	*Grains*	*Grams*	*Grains*	*Grams*	*Drams*	*Grams*
1/100	0.00065	2	0.130	26	1.70	3	11.65
1/64	0.00101	3	0.195	27	1.75	4	15.50
1/60	0.00108	4	0.260	28	1.82	5	19.40
1/50	0.00130	5	0.324	29	1.87	6	23.30
1/48	0.00135	6	0.400	30	1.95	7	27.20
1/40	0.00162	7	0.460	31	2.00	*Ounces*	
1/36	0.00180	8	0.520	32	2.10	1	31.10
1/32	0.00202	9	0.600	33	2.16	2	62.20
1/30	0.00216	10	0.650	34	2.20	3	93.30
1/25	0.00259	11	0.715	35	2.25	4	124.40
1/24	0.00270	12	0.780	36	2.30	5	155.50
1/20	0.00324	13	0.845	37	2.40	6	186.60
1/18	0.00360	14	0.907	38	2.47	7	217.70
1/16	0.00405	15	0.972	39	2.55	8	248.80
1/15	0.00432	15.5	1.000	40	2.60	9	280.00
1/12	0.00540	16	1.040	42	2.73	10	311.00
1/10	0.00648	17	1.102	44	2.80	11	342.14
1/8	0.00810	18	1.160	48	3.00	12	373.23
1/6	0.01080	19	1.240	50	3.25	14	435.50
1/5	0.01296	20	1.300	52	3.40	16	497.60
1/4	0.01620	21	1.360	56	3.65	24	746.40
1/3	0.02160	22	1.425	58	3.75	48	1492.80
1/2	0.03240	23	1.460	*Drams*		100	3110.40
3/4	0.04860	24	1.55	1	3.90		
1	0.065	25	1.62	2	7.80		

COMPARATIVE VALUES OF AVOIRDUPOIS AND METRIC WEIGHTS

Avoir. Ounces	*Grams*	*Avoir. Ounces*	*Grams*	*Avoir. Ounces*	*Grams*	*Avoir. Pounds*	*Grams*
$\frac{1}{16}$	1.772	5	141.75	13	368.54	4	1814.37
$\frac{1}{8}$	3.544	6	170.10	14	396.90	5	2267.96
$\frac{1}{4}$	7.088	7	198.45	15	425.25	6	2727.55
$\frac{1}{2}$	14.175	8	226.80	*Avoir. Pounds.*		7	3175.14
1	28.350	9	255.15			8	3628.74
2	56.700	10	283.50	1	453.60	9	4082.33
3	85.050	11	311.84	2	907.18	10	4535.92
4	113.400	12	340.20	2.2	1000.00		
				3	1360.78		

COMPARISON OF CENTIGRADE AND FAHRENHEIT THERMOMETRIC SCALES

CENT.	FAHR.	CENT.	FAHR.	CENT.	FAHR.	CENT.	FAHR.
100	212	62.2	144	24.4	76	13.3	8
98.9	210	61.1	142	23.3	74	14.4	6
97.8	208	60	140	22.2	72	15.6	4
96.7	206	58.9	138	21.1	70	16.7	2
95.6	204	57.8	136	20	68	17.8	0
94.4	202	56.7	134	18.9	66	18.9	2
93.3	200	55.6	132	17.8	64	20	4
92.2	198	54.4	130	16.7	62	21.1	6
91.1	196	53.3	128	15.6	60	22.2	8
90	194	52.2	126	14.4	58	23.3	10
88.9	192	51.1	124	13.3	56	24.4	12
87.8	190	50	122	12.2	54	25.6	14
86.7	188	48.9	120	11.1	52	26.7	16
85.6	186	47.8	118	10	50	27.8	18
84.4	184	46.7	116	8.9	48	28.9	20
83.3	182	45.6	114	7.8	46	30	22
82.2	180	44.4	112	6.7	44	31.1	24
81.1	178	43.3	110	5.6	42	32.2	26
80	176	42.2	108	4.4	40	33.3	28
78.9	174	41.1	106	3.3	38	34.4	30
77.8	172	40	104	2.2	36	35.6	32
76.7	170	38.9	102	1.1	34	36.7	34
75.6	168	37.8	100	0	32	37.8	36
74.4	166	36.7	98	1.1	30	38.9	38
73.3	164	35.6	96	2.2	28	40	40
72.2	162	34.4	94	3.3	26	41.1	42
71.1	160	33.3	92	4.4	24	42.2	44
70	158	32.2	90	5.6	22	43.3	46
68.9	156	31.1	88	6.7	20	44.4	48
67.8	154	30	86	7.8	18	45.6	50
66.7	152	28.9	84	8.9	16	46.7	52
65.6	150	27.8	82	10	14	47.8	54
64.4	148	26.7	80	11.1	12	48.9	56
63.3	146	25.6	78	12.2	10		

METHOD OF CONVERTING ONE SCALE INTO ANOTHER

1. To change a Fahrenheit reading into centigrade: Subtract 32 from the given degree (the freezing point in the Fahrenheit scale being this much higher than in the centigrade scale), multiply the remainder by 5, and divide the result by 9. Thus:

$$212° \text{F.} - 32 = 180 \times 5 = 900 \div 9 = 100° \text{C.}$$

2. To change a centigrade reading into Fahrenheit: Multiply the given degree by 9, divide the result by 5, and add 32 to the remainder. Thus:

$$100° \text{C.} \times 9 = 900 \div 5 = 180 + 32 = 212° \text{F.}$$

CHAPTER II

NATURE OF MATTER AND ITS ELEMENTS

Chemistry and Physics Defined—Nature of Matter, Molecules, Atoms—States of Matter—Molecular Motion—The Elements—Physical and Chemical Changes—Chemical and Physical Mixtures—Chemical Affinity—Cohesion—Adhesion.

Chemistry and physics defined.—Chemistry and physics are very intimately related since they are both concerned with the study of the nature of matter, especially inanimate matter, and with the changes which matter may be made to undergo. Chemistry, however, is more particularly devoted to the study of the composition of matter and is therefore often defined as *the study of the composition of matter*, and physics being more especially concerned with the properties and activities of matter has been defined as *the study of the properties of matter* and as *the study of matter in motion.*

Nature of matter.—By matter is meant *anything which possesses weight and which occupies space.* All matter is composed of inconceivably small particles called *molecules* and molecules are composed of still smaller particles called *atoms.*

Molecules.—A molecule is said to be *the smallest part of a compound that can exist by itself and main-*

tain its chemic features. Thus a molecule consists of two or more atoms. The atoms composing a molecule may be all of the same element, *e. g.*, all hydrogen or all oxygen, etc., or they may be of different elements, *e. g.*, a molecule of water consists of two atoms of hydrogen and one of oxygen.

Atoms.—An atom was formerly defined as *the smallest particle of an element that can exist* and atoms were thought to be absolutely indivisible, but it has been shown that atoms of elements are made of still smaller particles to which the name *electrons* has been given and a definition used for an electron is *a unit or atom of negative electricity.* Atoms do not, however, separate in ordinary chemical reactions and they are still generally regarded as being indivisible.

Compounds and elements.—A compound is a substance made up of two or more elements; an element is a substance which cannot be divided into simpler substances by any known means. For example, water is a compound of hydrogen and oxygen, but hydrogen and oxygen have never been subdivided and are therefore considered elements.

States of matter.—Matter is generally described as existing in three physical states, viz., solid, liquid and gaseous. All matter, however, cannot be classified as belonging strictly to any one of these three states and thus we speak of semi-solids and semi-liquids. Neither is the ordinary physical state of any kind of matter a fixed unchangeable thing, for by increase of temperature, solids can be liquefied and liquids vaporized (*i. e.*, made gaseous), and by putting gases under pressure in a low temperature, they can be changed to liquids and a low temperature will solidify liquids.

A very great change of temperature is not necessary to transform the physical state of some substances, but only very excessive changes will alter that of others.

One of the great differences in the nature of matter in its different physical states is the dissimilarity of the motion of its molecules.

Molecular motion.—It is the belief of scientists that the molecules composing all matter are constantly in rapid motion, but that differences exist in the kind of motion possessed by the different forms of matter. Thus, in solids, it is thought, the molecules oscillate with great rapidity about certain fixed points, but are held in their relative position by their attraction for each other or, as this is called in physics, *cohesive force*. However, even in solids, as will be discussed in the section on diffusion, the molecules sometimes break away. In liquids, though the molecules tend to cling together, they have no fixed position, but move about freely; thus liquids can adjust themselves readily to the shape of any vessel into which they are put, and when different liquids are poured into a vessel, they soon become mixed, as they would not do if their molecules were fixed and motionless. The movement of gas molecules is very marked and they exert no cohesive force whatever upon each other, but, on the contrary, they tend to separate and fly apart indefinitely, spreading in all directions. One evidence of this is the rapidity with which the odor of a gas (*e. g.* ammonia) will be perceived at a distance from an open vessel containing even a solution of the gas. The odor of the gas could not become thus diffused if the molecules were stationary in the vessel for there would be no odor were there not some molecules of the gas present in the air that is inhaled.

Until the idea of molecular motion was conceived, scientists found it impossible to explain many common phenomena, such as pressure of gases, evaporation, contraction, and expansion.

Number and condition of elements.—There are about eighty-three elements known at the present time. Of these, eleven are ordinarily gaseous—the principal gaseous ones being oxygen, hydrogen, nitrogen, and chlorin—two are liquid—mercury and bromin, and the others are solid.

Symbols.—In order to facilitate the explanation of the composition of chemicals and of chemical reactions, chemists have adopted symbols to represent the elements. These symbols consist of the initial letter of either the Latin or the English name of the element, or when there are two or more elements having the same initial letter, of the initial and one other letter. For example, C stands for carbon, Cl for chlorin, N for nitrogen and Na for sodium, the Latin name of which is *natrium*. A number placed after a symbol shows that there are as many atoms of the element as the number designates; thus H_2O (water) signifies that in each molecule of water there are two atoms of hydrogen and one atom of oxygen.

Atomic weight.—Atoms of the elements are believed to have fixed weights, but, as it would take about 200,000,000 hydrogen atoms placed side by side to form a line one centimeter in length, their weight can be calculated only as relative to some fixed standard that is given a supposititious weight. For example, hydrogen, which until lately was the lightest element known, was used as the standard and given a weight of one microcrith,[1] and the weight of each of the other elements was determined by finding the weight of the amount that combined with a definite weight of hydrogen. Lately, oxygen has been chosen as a standard because as more elements unite to form definite compounds with it than with hydrogen it is easier to compare their weight with it than with the latter. The atomic weight of oxygen, however, is sixteen microcriths.

Molecular weight.—In studying chemistry, one hears constantly not only of atomic weights, but also of molecular weight. By the weight of a molecule is meant the sum total of the weights of all the atoms composing it. Thus, the molecular weight of calcium carbonate ($CaCO_3$) will be 100 because, as shown by the symbols, it is a compound consisting of 1 atom of calcium, which weighs 40, one atom of carbon, which weighs 12, and three atoms of oxygen each one of which weighs 16; thus $40+12+48=100$.

The following table gives the names of the elements, their derivation and atomic weights. It is of course not necessary to memorize any of the weights, and only the elements marked with an asterisk need be remembered.

A TABLE OF THE ELEMENTS

ELEMENT	SYMBOL	ATOMIC WEIGHT	NATURE	WHERE FOUND IN NATURE
*Aluminium (Aluminum)	Al	27.1	A white metal that is very light, stiff, and strong. It is a good conductor of heat and of electricity.	In many soils and rocks, where it exists in the form of oxids and salts.

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[1] The microcrith is defined as *the weight of one atom of hydrogen*. It is not any definite weight and is so infinitely small that it cannot be compared with such weights as grams and grains.

Element	Symbol	Atomic Weight	Nature	Where Found in Nature
Antimony (Lat. *Stibium*)	Sb	120.2	A bluish-white, hard, brittle, solid.	Occurs in nature chiefly as a sulphid called stibnite, but is found also as an oxid and as a constituent of various complex minerals.
Argon	A	39.9	Gas.	In the air.
Arsenic	As	75.0	A steel-gray brittle metallic-looking substance.	In various ores. It occurs most frequently in the form of salts and other compounds the majority of which, unlike pure arsenic, are highly poisonous.
*Barium	Ba	137.4	A pale-yellow metal.	In minerals.
Beryllium	Be	9.1	Metal.	In beryl and in various minerals.
*Bismuth	Bi	208.5	A grayish-white crystalline solid.	Occurs in certain localities in uncombined state, and also as an oxid and sulphid.
Boron	B	11.0	Resembles silicon.	Chiefly in borax and boracic acid.
*Bromin	Br	79.96	A dark red liquid about three times as heavy as water. Its vapor has an offensive odor and is very irritating to the eyes, throat, etc.	In sea salt and the salt of various springs and salt deposits.
Cadmium	Cd	112.4	Resembles zinc.	In zinc ores.
Cæsium	Cs	132.9	A white metal.	Chiefly in mineral springs.
*Calcium	Ca	40.1	Silvery-white metal.	In various compounds as marble, limestone, fluorspar, phosphorite, gypsum, etc.
*Carbon	C	12.00	A solid that is amorphous in charcoal; crystalline in diamond and graphite. Its compounds are numerous and occur as gases, liquids and solids.	In all organic matter and a few inorganic substances as diamonds, calcium carbonate, sodium carbonate, etc.
Cerium	Ce	140.25	Resembles iron.	In a few rare minerals.
Chlorin	Cl	35.45	A greenish-yellow gas. It has a peculiar suffocating odor, and, if inhaled, produces a severe irritation of the respiratory tract. It has a strong affinity for hydrogen and it combines readily with many metals.	In sea-water and in salt-deposits in the forms of chlorids, especially those of sodium, potassium, and magnesium. It is never found free in nature.
Chromium	Cr	52.1	A hard metal.	In chromite or chrome iron ore.

ELEMENT	SYMBOL	ATOMIC WEIGHT	NATURE	WHERE FOUND IN NATURE
Cobalt	Co	59.0	A reddish or white metal (some of its salts are blue).	Occurs chiefly in combination with arsenic and sulphur.
Columbium	Cb	94.0	Metal.	In columbite and a few other rare minerals.
*Copper (Lat. *Cuprum*)	Cu	63.6	A reddish colored metal.	Free and in various ores.
Coronium		Not known	Gas.	In the sun.
Crypton	K	80.0	A very inert gas.	In the air.
Erbium	Er	166.0	A metal-like solid.	In some rare metals,
Fluorin	F	19.0	Gas, resembles chlorin.	In a few minerals, chiefly fluorspar.
Gadolinium	Gd	156.0	Solid.	In a few minerals.
Gallium	Ga	70.0	Metal.	In some zinc blends.
Germanium	Ge	72.5	Metal.	In the metal argyrodite.
Glucinum	Gl	9.03	Metal.	In beryl and a few minerals.
*Gold (*Aurum*)	Au	197.2	Metal.	Free and in ores.
Helium	He	4.0	A light gas.	In the atmosphere.
*Hydrogen	H	1.008	A very light gas.	In water and in many other compounds both organic and inorganic.
Indium	In	113.6	Metal.	In zinc ores.
*Iodin	I	126.97	Purplish-black crystals which vaporize at ordinary temperatures, giving off a violet-vapor of unpleasant odor.	Iodin occurs free in sea-water, from which it is absorbed by certain sea-plants, so that it is found in their ashes. Also it is found in salt springs and beds and in Chili saltpeter.
Iridium	Ir	193.0	Metal.	In a rare mineral called iridosmin.
*Iron (*Ferrum*)	Fe	55.9	Metal.	A common constituent of rocks and soils. It is assimilated by plants and animals and is essential to their life.
Krypton	Kr	81.8	Gas.	In the atmosphere.
Lanthanum	La	138.9	Metal.	In a few rare metals.
*Lead (*Plumbum*)	Pb	206.9	Metal.	Occurs chiefly as a sulphid called *galena*.
Lithium	Li	7.03	Silvery-white alkali metal.	In a few metals, as lepidolite, and, in the form of carbonates and chlorids, in some mineral waters.
*Magnesium	Mg	24.36	White metal.	It is a constituent of many rocks, also salts of the metal are found in sea-water and salt-deposits.
*Manganese	Mn	55.0	A hard gray metal somewhat like iron.	In some iron ores and minerals and as a dioxid called *pyrolusite*.

ELEMENT	SYMBOL	ATOMIC WEIGHT	NATURE	WHERE FOUND IN NATURE
*Mercury (*Hydrargyrum*)	Hg	200.00	A heavy silvery liquid.	It occurs chiefly as a sulphid called *cinnabar*, and in globules of metal inclosed in the cinnabar.
Molybdenum	Mo	96.0	A metal-like element.	In molybdenite.
Neodymium	Nd	143.6		In a few rare minerals.
Neon	Ne	20.0	An inert gas.	In the air.
Nickel	Ni	58.7	Metal.	In metallic ores. It is usually found combined with arsenic or sulphur.
*Nitrogen	N	14.04	Gas.	In the air and in many organic and inorganic substances. It is an essential constituent of all living matter.
Osmium	Os	191.0	Metal.	In platinum and iridosmin.
*Oxygen	O	16.00	Colorless, odorless, tasteless gas, slightly heavier than air.	It occurs free in the air; in the combined state, it forms $\frac{8}{9}$ of water and enters into the composition of all organic, and nearly all inorganic, compounds.
Palladium	Pd	106.5	A silver-white metal.	Free and in combination with other metals, especially gold and platinum.
*Phosphorus	P	31.0	A yellowish, waxy, soft solid.	In the form of various phosphates. P. occurs in many organic and inorganic substances.
*Platinum	Pt	194.8	A grayish-white metal of high luster, very malleable and ductile.	Chiefly alloyed with gold and what are known as the platinum metals, *e. g.*, rhodium, pallidium, and osmium.
*Potassium (*Kalium*)	K	39.15	A soft alkali metal. It decomposes water very vigorously and the heat evolved in the chemical reaction is sufficient to ignite the hydrogen set free.	In many rocks, minerals, and organic compounds. It is never found free in nature.
Praseodymium	Pr	143.5		In a few minerals.
Radium	Ra	225.0	. . .	In pitchblend.
Rhodium	Rh	102.9	White metal.	Usually in connection with platinum.
Rubidium	Rb	85.5	Metal.	In certain mineral springs.
Ruthenium	Ru	101.7	Metal.	Associated with platinum.
Samarium	Sm	150.3		In a few rare minerals.
Scandium	Sc	43.97	. . .	In a few minerals.

ELEMENT	SYMBOL	ATOMIC WEIGHT	NATURE	WHERE FOUND IN NATURE
Selenium	Se	78.87	A non-metallic solid.	In some minerals and sulphur.
*Silicon	Si	28.4	A non-metallic solid substance that occurs in both crystalline and amorphous forms.	Next to oxygen, silicon is the most abundant element in nature; it is never found free, but its compounds constitute a large portion of the earth's crust. It is generally prepared from quartz.
*Silver (Lat. *Argentum*)	Ag	107.66	A lustrous white metal.	Free and in many ores.
*Sodium (Lat. *Natrium*)	Na	23.05	A soft alkali metal that resembles potassium.	It is never found free, but its compounds are numerous and widely distributed. They occur in rocks, sea-water, salt deposits, and organic matter.
Strontium	Sr	87.6	An alkaline earth.	In the metals strontianite and celestite.
*Sulphur	S	32.06	A pale-yellow crystalline solid.	Free, especially in the neighborhood of volcanoes, and in the form of sulphates and sulphids in a variety of minerals and many organic substances.
Tantalum	Ta	183.0	A black powder-like solid.	In a few minerals.
Tellurium	Te	125.0	A silver-colored crystalline solid.	In a few minerals.
Terbium	Tb	159.0	. . .	In a few minerals.
Thallium	Tl	204.1	A soft white metal belonging to the same class as aluminium.	In many ores, in the minerals known as *pyrites*, and in combination with sulphur.
Thorium	Th	231.9	A gray metallic powder of the nature of cerium. (Oxids of these two elements are used in the preparation of Welsbach mantles, because of the intense light given out by a mixture of these oxids when they are heated.)	In a few rare minerals.
Thulium	Tm	171.0	. . .	In a few rare minerals.
*Tin (*Stannum*)	Sn	119.0	A soft metal that melts at about 235° F. and is very malleable.	It occurs chiefly as the oxid called *cassiterite* or *tinstone*.
Titanium	Ti	48.1	Resembles silicon.	In soil and rocks.
Tungsten (Wolframium)	W	184.0	A very hard, brittle, nearly infusible metal.	Chiefly as ferrous tungstate in the mineral wolfram.

ELEMENT	SYMBOL	ATOMIC WEIGHT	NATURE	WHERE FOUND IN NATURE
Uranium	U	238.5	A rare metal belonging to the same class as tungsten.	In pitchblend and a few minerals.
Vandium	V	51.2	A crystalline metallic substance.	In a few minerals.
Victorium	Vi		Metal.	In earth.
Xenon	Xe	128.0	Gas.	In the air.
Ytterbium	Yb	172.6	A solid substance.	In a few minerals.
Yttrium	Yt	89.0	An earth metal.	In a few minerals.
*Zinc	Zn	65.4	A heavy bluish metal.	In a few ores.
Zirconium	Zr	90.6	A solid that occurs in both crystalline and amorphous forms.	In a few rare minerals.

Physical and chemical changes.—Everywhere in nature change is constantly taking place. Matter is being continuously broken down, complex substances reduced to simple ones, and these simple substances are as constantly being built up, sometimes into very different combinations from their original form. All life and growth are dependent upon such changes. According to their nature these changes are classed as physical and chemical; physical changes being those which do not involve a necessarily permanent change in a substance and chemical changes those which are permanent or of a kind that cannot be readjusted without chemical means. For example, when cane sugar is dissolved in water, a physical change has taken place, because if the water is evaporated, the sugar will be obtained in the same form as it was originally, but, if cane sugar is boiled with acid, it will be changed to quite a different substance, viz., glucose; this, therefore, is a chemical change. The freezing and vaporization of water are physical changes, because if the ice is melted or the steam condensed, water, similar to that frozen or vaporized, will be obtained. If a substance (*e. g.*, a piece of wood) is burned, a chemical change is perpetrated, for by no known means can the ashes which remain and the gases which have passed into space be put together again to form a piece of wood like unto that which was burned.

Physical or mechanical mixtures and chemical compounds.—When two or more substances of a different nature are mixed together, but do not enter into direct combination with, or change the nature of, each other, the mass is called a *mechanical mixture*, but when the substances combine, forming matter of a different nature, the result is known as a *chemical compound*.

Chemical affinity. Cohesion. Adhesion.—One cause for both the stability and change of matter is chemical affinity. By this is meant *the electrical attraction between the different atoms which causes elements to combine and which holds them together when combined.* This will be further discussed in Chapter X. In physics, the attraction which holds molecules of a like nature together is spoken of as *cohesion* and that which holds substances of unlike character together is called *adhesion*.

CHAPTER III

ENERGY, HEAT, AND PRESSURE

Definition of Hypothesis, Theory, Law—What is Meant by Energy—Effects of Heat—Different Ways in which Heat is Transmitted—Difference between Amount and Degree of Heat—Heat Units—Specific Heat—Latent Heat—Heat of Fusion—Relation of Pressure and Heat—Atmospheric Pressure—The Effect of Pressure upon the Boiling Point of Liquids—The Effect of the Specific Gravity of Liquids upon their Boiling Point.

The changes that occur in nature, either chemical or physical, are not carried on in a haphazard manner, but are controlled by absolute, often unalterable, forces. In referring to or describing these forces of nature, the scientist often speaks of them as *laws*. There are, in fact, three words often used by scientists in this connection, viz., *hypothesis*, *theory*, *law*.

Definitions of above terms.—An hypothesis is sometimes defined as *a suggestion, the truth of which has not been proved, that is advanced as a reason for, or as a cause of, some fact or law*. An hypothesis is also often spoken of as a *working theory*, because an unproved suggestion or idea of what may be the cause for certain effects—in other words an hypothesis—is often the first step in the work of discovering hidden forces of nature or of the application of these forces to practical inventions. By theory is meant a long-standing hypothesis for phenomena that accounts *for them so satisfactorily that it seems to be true, though its validity has never been proved.* When a theory has been tested in every known way and is proven to be invariably accurate it is called a *law.*

The following laws and hypotheses, which are a few of those frequently referred to in physics and chemistry, will be often alluded to in this or following chapters.

Atomic hypotheses.—(1) All elements are made up of minute, independent particles called *atoms.* (2) All atoms of the same kind of element are alike, atoms of different kinds of elements are different. (3) All atoms combine by wholes to form compounds; atoms are never divided in chemical reactions.

Law of definite composition.—The composition of chemical compounds never varies; *e. g.*, water, whether it exists in liquid, solid, or gaseous form, consists of H_2O.

Law of the conservation of energy.—Energy can be changed from one form to another, but it cannot be either created or destroyed. For example, radiant energy from the sun produces chemical energy in plants which causes them to grow. This energy remains stored in plants in what is known as *potential* or *latent* energy. If the plants are used as food, this stored energy, after the food has been digested, absorbed, carried to the tissues, and there decomposed by oxidation, will be set free and will appear in the form of heat or of kinetic energy, *i. e.*, the energy of motion. Or, for another example, the plants that have decayed and in the course of time turned to coal, contain, stored in their substance, the energy, derived from the sun, which furthered their growth, and when the coal is set on fire this stored energy is transformed into heat, light, or electrical energy; and though energy thus used may seem to have come to an end, it has simply assumed some other form and exists somewhere. Thus the amount of energy in the universe remains unchanged.

Law of the conservation of matter.—Matter can be changed from one form to another, but it cannot be either created or destroyed. For example, a house may be burned, but the elements composing it are still in existence, some in the form of ashes, some as gases that have passed into space. The oxygen of the air, which maintained the combustion, united with some of the carbon and hydrogen in the burning matter, forming thereby carbon dioxid (CO_2) and water (H_2O); this will be absorbed by plants and by them changed to the substances they need for their growth, and while doing this they will give off oxygen and thus the oxygen will be returned to the air once more. If the plants die, the elements of which they are composed may be scattered, but they will still exist.

Laws of boiling.—(1) Under a given pressure, every liquid has a definite boiling point; for example water boils at 100° C., alcohol boils at 80° C., mercury boils at 350° C. (2) When the boiling point is reached the temperature remains constant unless the pressure is increased.

Law of Boyle.—The volume of a gas varies with the pressure that is put upon it; *i. e.*, when a gas is under heavy pressure, it will occupy but a small space, but when pressure is released, the gas will expand and thus increase its volume.

Law of Charles.—Every true gas expands $\frac{1}{273}$ of its volume for each degree centigrade that its temperature is increased and it contracts $\frac{1}{273}$ of its volume for each degree centigrade that its temperature falls.

Laws of liquid pressure.—(1) Pressure in liquids is proportional to the depth alone and is not influenced by the size or shape of the vessels which contain them. (2) At any given depth the pressure is equal in all directions.

Laws of pressure in gases.—(1) Pressure in gases increases with depth, but is not proportional to it. (2) At any given depth the pressure is equal in all directions.

Difference between the use of the terms gas and vapor.—The word *gas* is usually applied to matter like nitrogen, oxygen, and hydrogen which can be liquefied only under high pressure, and the term *vapor* is used to express the gaseous state of matter that is liquid under ordinary conditions, *e. g.*, water, alcohol, ether. The term *aqueous vapor* is often applied to vapor derived from water. Vapors and gases are governed by the same laws.

Energy

By energy is meant *the power to produce motion*, or, in other words, *the power to perform work.* Energy manifests itself in several forms,—for example, we have heat, light, mechanical energy, electrical energy, and chemical energy,—and one form can be made to produce another form, or, in the words of the law of the conservation of energy, *energy can be changed from one form to another.* Heat, for example, or chemical action can be made to produce an electric current and an electric current will produce heat and promote chemical activity in matter.

When a body is at work its energy is spoken of as kinetic energy (from the Greek *kinema*—motion). When a substance containing the power of work within itself or of causing motion in other matter is at rest or inactive, its power is spoken of as *latent or potential energy.*

A body possesses energy only because it has received energy in some form and it possesses only as much as it has received. Nearly all the available energy upon earth is derived primarily from the sun. The most obvious example of the sun as a source of energy is that given under the law of the conservation of energy. As there stated, the sun provides the energy which causes plants to grow, and plants pass this energy on to man and other animals who use plants as food, or when substances derived from plants, as wood, coal, etc., are burned the energy they liberate is used to run machinery, produce electric currents or perform other work. Other examples of the sun as a source of energy will be seen in the sections on radiation and reflection.

Heat

As can be seen in the laws and hypotheses on page 39, heat and pressure are two important factors in causing changes in matter.

Definition.—Physics and physiology give somewhat different definitions for the term heat. A definition given in physics is *the molecular motion of a body*, for heat is due to the vibration of the molecules of matter, and the greater the degree of vibration, the more intense the heat, and, conversely, the more intense the heat, the more rapid the motion. It is because heat increases molecular motion that it changes the state of matter, for when molecular motion becomes very rapid, it overcomes the cohesive force of the molecules which hold them together. A definition for heat given in physiology is *the sensation perceived upon contact with hot matter due to the stimulation of certain nerve endings by the rapidly moving molecules of matter.*

Sources of heat.—These are (1) the sun; (2) chemical energy, *e. g.*, the heat produced (a) by combustion, (b) by the oxidation that goes on within the body and other forms of slow oxidation, (c) the heat evolved during certain chemical reactions, as the slaking of lime; (3) electrical energy; (4) mechanical energy, *e. g.*, (a) friction (machinery, for instance, becomes hot while working when its parts rub upon one another, thus causing friction), (b) percussion (*e. g.*, when a nail is hammered, both nail and hammer become hot); (5) pressure, see page 48.

Experiment 4. Object, to show the rapidity with which heat can be produced by chemical reactions.

(a) Add some sulphuric acid to about four drams of cold water and keep a thermometer in the beaker while doing so and for a short time afterward. Note the temperature.

(b) Add water to some quicklime and note the temperature while the lime is slaking.

Effects of heat.—The more important effects of heat are: (1) Change of size—heat expands matter; (2) change of state—*e. g.*, ice is melted, water is vaporized; (3) change of temperature; (4) chemical change—*e. g.*, sugar is changed to caramel, glucose, etc.; (5) electric change—*e. g.*, the production of an electric current by the heating of two different metals at their junction.

How heat is transmitted.—The way in which heat is transmitted from one place or body to another depends upon the form of matter transferring the heat. Thus heat passes through solids by the process known as conduction; it is transmitted through liquids, mainly, by the process of convection, and through the ether by radiation.

Difference between the amount and degree of heat.—The amount of heat in a body and the degree of heat of a body, or in other words the temperature of a body, are not dependent upon each other; *e. g.*, the temperature of 8 ounces of boiling water will be the same as that of 2000 ounces of boiling water, but there will be a much larger amount of heat in the vessel containing the 2000 ounces than in that holding the 8 ounces. The degree of heat of a body is ascertained by the thermometer and is recorded in degrees; the amount of heat in matter is ascertained by the use of a calorimeter and recorded in calories—a calory being the amount of heat necessary to raise 1 gram of water 1 degree centigrade, or what is known as the large calory, 1 kilogram of water 1 degree centigrade. The calory is spoken of as the *heat unit.* In Great Britain the amount of heat is often recorded in what is called the *British Thermal Unit* or B. T. U. By a B. T. U. is understood the amount of heat required to raise the temperature of 1 pound of water 1 degree Fahrenheit. A large calory is about the equivalent of 4 B. T. U.

Specific heat.—By this is meant *the amount of heat which is required to raise the temperature of a given amount of any substance 1 degree centigrade* (or any other degree taken as a standard) *as compared with that required to raise an equal quantity of water to the* *same temperature.* It requires 1 calory of heat to raise the temperature of 1 gram of water 1 degree centigrade; therefore the specific heat of water is said to be 1. The specific heat of the majority of substances is less than that of water. As a rule the denser a substance is, the lower its specific heat; *i. e.*, it takes less heat to raise its temperature. Thus:

The specific heat of gold is 0.032
" " " copper is 0.093
" " " glass is 0.19
" " " earth is 0.2
" " " air is 0.237
" " " steam is 0.480
" " " ice is 0.504
" " " alcohol is 0.604
" " " water is 1.
" " " hydrogen is 3.409

In some connections, the specific heat of substances is often spoken of as their *heat capacity* as it practically signifies the capacity that they have for absorbing heat.

Value of high heat capacity of water.—The large amount of heat that water can absorb and the slowness with which it parts with its absorbed heat is of great value to man. For instance, the land by the seashore becomes hot quicker and to a greater degree than the water while the sun is shining upon them, but the land parts with its heat more rapidly and becomes cooler than the water during the night. This, as will be seen in the section on convection, gives rise to the breezes that constitute one reason for the more even temperature and the greater coolness in summer of countries near large bodies of water than those of the same latitude in the interior. Another reason for this is that, during the summer, the water of the oceans and large rivers and lakes, etc., is constantly absorbing heat which in winter unless frozen it slowly parts with.

Common uses to which the capacity of water to absorb and part with a large amount of heat is put is the use of steam and hot water for heating houses, the use of hot-water bags, and the like.

Relative heat of bodies.—The relative heat of different kinds of matter cannot be always judged by touching substances, because the degree of the sensation of heat experienced depends not only upon the degree of heat in the substance, but also upon the speed with which it absorbs and parts with heat when in contact with the hand; *e. g.*, if a piece of iron and a piece of wood are heated to the same temperature, the iron will seem hotter to the touch than the wood because it gives up its heat more readily, but if the iron and wood are both equally cold, the iron will feel colder than the wood, because iron absorbs heat more rapidly than wood and will take heat from the hand.

Latent heat. The heat of fusion.—The examples usually used to explain latent heat are the liquefaction of solids and the vaporization of liquids. When pieces of ice are heated their temperature rises until 0° C. (32° F.) is reached, then the ice begins to melt and its temperature does not increase until the melting is accomplished because the heat is used to overcome the cohesion of the molecules of the matter and to move them apart, this being necessary for melting. Heat absorbed in this fashion without causing rise of temperature is said to be *latent* and that used for the liquefaction of solids is often spoken of as the *heat of fusion.*

Heat is always necessary for the liquefaction of solids, but the amount of heat required varies with different kinds of solids. It requires 80 calories of heat to melt 1 gram of ice at 0° C. without raising the temperature.

The latent heat of steam.—As stated under the laws of boiling, when water reaches a temperature of 100° C. (212° F.) the temperature ceases to rise unless the steam is put under pressure. One reason for this will be found on page 51; another reason is that at 100° C. water is converted into steam—*i. e.*, vaporized—and heat is required to further the process. It requires 536 calories of heat to change 1 gram of water into steam. In other words a gram of water must absorb the amount of heat reckoned as 536 calories, or a liter of water 537,000 calories, before it will be vaporized.

Why the heat used for the fusion of solids and the vaporization of liquids is said to be latent.—When the term latent was first applied to the heat used to cause fusion or vaporization, heat was thought to be a weightless fluid that disappeared in matter during liquefaction or vaporization and passed out once more when the liquid solidified or the vapor condensed. It is now known, however, that not only was this conception of heat wrong, but that, though the heat is not raising the temperature of the water, it is not quiescent, for it is at work increasing the molecular motion of the matter, thus breaking the cohesion that holds the molecules together, causing them to separate and keeping them apart. It will be remembered that cohesive force is stronger in solids than in liquids and that it does not exist in gases. As heat used in liquefying and vaporizing is doing work, the terms *heat of fusion* and *heat of vaporization* are now considered more accurate than the expression *latent heat*, though this latter term is still much used.

Recovery of the heat of fusion, etc.—Just as much heat as is used in matter in promoting the molecular motion that causes the changes described in the preceding paragraph passes from the matter as it resumes its original form; thus, for every gram of water that freezes, 80 calories are liberated, and for every gram of steam that is condensed, 536 calories are set free.

How water, cream, etc., are cooled by melting ice.—When ice is put into a pitcher of water, it cools the latter not only because the ice and the water from it are colder than the water in the pitcher, but because the ice takes the heat it needs for melting from the water. In making ice cream, etc., the cream is frozen because the ice takes the heat it needs for melting from the freezer and its contents. Salt is added to the ice because it hastens the freezing in several ways: (1) It hastens the melting of the ice, because the salt molecules attract the ice molecules and this helps in the separation of molecules that is essential for liquefaction, and the quicker liquefaction occurs, the more rapidly will heat be taken. (2) The salt, being soluble in water, absorbs the latter and takes the heat it requires for the process from the cream, etc. (3) The lowest temperature of unfrozen water is 32° F., that of an unfrozen, saturated salt solution is about 6° F.; thus the brine produced as the ice and salt melt is much colder than the water of melted ice would be.

The greater the amount of salt used, the quicker the cream will be frozen, but if it is frozen very quickly, it will be of a coarse, granular texture. The usual proportion of salt to ice used is one part salt to three of ice.

Some other practical applications to which the requirement of heat for fusion and vaporization are put will be given in the sections on evaporation and condensation.

Pressure in Gases

How pressure is increased.—To understand how gases and vapors can be made to produce increase of pressure, it is necessary to realize that the molecules of which matter is composed are not, even in solids, packed excessively tightly together and that the tightness with which they are packed is less in liquids than in solids, while gas molecules tend to separate and fly apart. Another point to remember is that molecular motion is increased by heat and lessened by cold.

Bearing these things in mind, it can be realized that when a gas is restricted within a given space, the moving molecules strike and cause pressure against each other and against the walls of the vessel confining them. If the space is large in comparison to the amount of gas present the pressure will be slight, but if, by any means, the size of the space is lessened, the moving molecules will come more in contact with each other and the sides of the vessel and thus produce increased friction. If the gas is heated, the movement and consequently the friction and pressure will be still further increased, and the friction will increase the heat. Thus pressure produces heat and heat produces pressure

Comparative relation of pressure and heat.—When steam, for example, is put under pressure, as in an autoclave, there is a definite relationship between the amount of pressure exerted and the degree of heat obtained. The amount of pressure exerted is usually expressed as being so many pounds to the square inch of surface exposed to the pressure. The approximate comparative relation of the amount of pressure and degree of heat can be seen in the following table[1]:

Pounds pressure to the square inch	Temperature
1	102° C.
5	108° "
10	115.6° "
15	121.3° "
20	126.2° "
30	134.6° "

In some strongly constructed boilers a pressure of even 200 pounds to the square inch can be obtained; in fact, so great is the pressure exerted by steam, even exceedingly strong boilers can be shattered by allowing the temperature to increase, and thus raising the pressure, above the degree that the boiler was constructed to stand.

Atmospheric pressure.—Since the pressure of gases increases with depth, the pressure of the atmosphere is greater at sea level than in high mountains. At a height of three miles, the pressure is only about one half that exerted at sea level. Below sea-level, the air increases rapidly in density and, therefore, it naturally exerts much greater pressure. It is estimated that thirty-five miles below the sea the pressure must be a thousand times greater than that at the earth's surface.

The barometer.—Torricelli, an Italian physicist, was the first individual to measure atmospheric pressure.

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[1] It will be remembered that steam not under pressure has the same temperature as boiling water, *i. e.*, 100° C., or 212° F.

He took a glass tube, about one meter in length and five millimeters in diameter, closed it at one end, and completely filled it with mercury. Then, putting his finger over the opening, he inverted that end of the tube into a pan containing mercury. Then he withdrew his finger and found that the mercury remained at a considerable height in the tube. This, by other experiments, he proved to be due to the pressure of the air upon the surface of the mercury in the dish.

A common form of barometer is now constructed on the same principle as the Torricellian experiment. It consists of a mounted tube at the lower end of which is a bulb or crook of exceedingly thin glass for the mercury reservoir. The pressure of air upon this causes the mercury to rise 760 mm. in the tube when the barometer is at sea level and the temperature at zero centigrade (*i. e.*, 32° F.). Calculations that need not be given here show that if the mercury is raised to 760 mm. (or as it is sometimes expressed 76 cm.) in the tube, the atmospheric pressure is fifteen pounds to the square inch at sea level, when the temperature is zero centigrade.

Common causes other than altitude that produce changes in atmospheric pressure.—Disturbances in the atmosphere, such as those produced by variations in temperature, winds, and humidity alter the degree of pressure in a locality. Heat reduces pressure because it causes increased movement and expansion of the air, thus making it less dense. When pressure is thus reduced, currents of air from cooler regions pass in; if they do so quickly, there is a wind; if there is much humidity in the air, the cool currents will cause the vapor to condense and there will be rain or snow, etc. When the atmospheric pressure is low, the mercury in the barometer falls; thus it can be seen why a *falling mercury* is looked upon as presaging a storm.

Effect of pressure upon boiling point. Nature of boiling.—A liquid is said to be boiling when bubbles of vapor form within it. Bubbles cannot form in a liquid while the atmospheric pressure at the surface is greater than that of the vapor as there is no room for their expansion, but the pressure of the vapor that is formed from the water increases as its temperature rises because its molecular motion increases and when a temperature of 100° C. (212° F.) is reached, if the liquid is water, the pressure of the vapor is as great as that of the air on the surface and the bubbles rise to the top. After a liquid begins to bubble its temperature cannot rise higher unless the steam is put under pressure for the reason given on page 46, and because the bubbles increase the surface at which evaporation takes place and this causes loss of heat. Consequently, to make water boil violently is simply a waste of fuel, for the larger the bubbles, the greater the loss of heat by evaporation and thus there is no rise of temperature. The temperature of flowing steam (*i. e.*, that not under pressure) is the same as that of boiling water.

Experiment 5. Object, to show the effect of atmospheric pressure upon boiling point.

Articles required: A flask about half full of water, a cork with one hole in it—the cork must fit the bottle perfectly,—a thermometer that will fit into the hole in the cork, a Bunsen burner, an iron stand, wire netting, a towel, and, unless the cold water faucet is in a good light, a pitcher of cold water and a basin.

Procedure: Boil the water in the flask for about ten minutes, then put the cork, with the thermometer in the hole (the thermometer must extend into the water), into the flask, and *at once* turn out the light, grasp the upper part of the neck of the flask with the folded towel and run cold water over the upper part of the flask, above, not over, the part containing the water. See Fig. 39. Watch the thermometer.

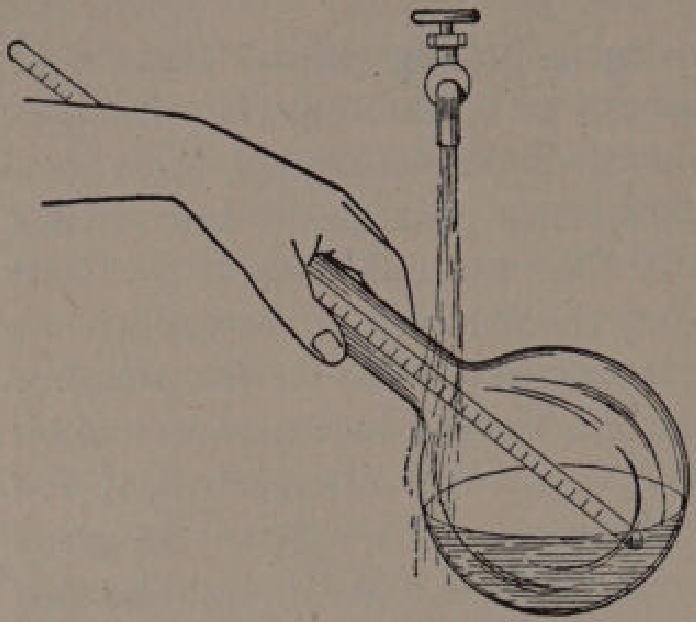

FIG. 39. MANNER OF POURING COLD WATER RUN OVER FLASK IN ORDER TO REDUCE THE PRESSURE ABOVE THE WATER WITHIN IT.

What happens?

Why?

Endeavor to answer these questions before reading the following paragraph.

Explanation of experiment. The air in the flask was heated, as well as the water, while the latter was boiling, therefore it expanded and, consequently, passed from the flask and its place was taken by the steam produced as the water boiled. When the flask was held under the cold water, the steam was condensed and thus there was a partial vacuum above the water; this being the case, the pressure was, of course, reduced and the vapor, having less pressure to overcome, bubbled at a lower temperature.

It is because the pressure on high mountains is less than at sea level that water boils there at a lower temperature than it does in lower altitudes. At the top of Mount Blanc, water boils at 84° C.

Why pressure makes it possible for some bodies to float.—Due to the fact that the pressure of liquids is equal in all directions, upwards as well as downwards, any body that does not weigh more than the amount of liquid which it displaces will float since its weight is not more than that of the amount of liquid that would otherwise have been in that part.

The nature of specific gravity or density and its influence on floating bodies.—First what is specific gravity or density? These two words are used interchangeably to signify the weight of a substance as compared with that of another substance which is taken as a standard. Distilled water is the standard usually employed, especially for liquids. A liquid in which solids heavier than water are in solution or suspension will, of course, have a higher specific gravity than water. Naturally, the greater the density of the liquid, the heavier the body that will be able to float in it. It is impossible for people to sink in the Great Salt Lake of Utah.

How the specific gravity of liquids is ascertained.—The specific gravity of such liquids as water, milk, sugar solutions, acids, urine, blood, and the like is usually gaged with a hydrometer. This consists of a glass cylinder on one end of which is a bulb containing mercury or shot and on the other a slender stem within which is a scale (see Fig. 23). The scale varies slightly in different forms of hydroeterms. The scale of an instrument intended for measuring the density of liquids lighter than water is usually marked in such a manner that when the instrument is put into distilled water, it sinks until the mark 1 is on a line with the water, and when put into pure alcohol, it sinks until the mark 100 is reached. The scale of another variety of hydrometer is marked so that the instrument will sink to 1 in distilled water and the markings above the 1 indicate weights less than 1 and those below that point show densities greater than that of water. Thus the instrument will not sink to as great an extent in milk, blood, and other dense liquids as in water. The principle involved is that *a floating body sinks until it displaces its own weight.* Determination of the sp. g. of a liquid is of value in testing for its purity.

Experiment 6. Object, to show the effect of specific gravity on the boiling point of liquids.

Procedure: (a) Ascertain the specific gravity of tap water, boil, and notice the temperature at which boiling occurs. (b) Add three tablespoonfuls of salt to the water, cool, ascertain the sp.g., boil, and note the temperature at which boiling occurs. (c) Take equal parts of water and alcohol, find the sp.g., boil, and note the temperature.

From this experiment it can be realized that liquids with a lower sp.g. than water boil before the temperature reaches 100° C., while those of greater density than water have a higher boiling point than the latter. For example:

The boiling point of ether is...........	37°C.
" " " " alcohol 95%.......	79 "
" " " " water.............	100 "
" " " " spts. of turpentine..	130 "
" " " " phosphorus........	290 "
" " " " mercury..........	357 "

CHAPTER IV

SOME COMMON PHYSICAL PROCESSES AND THEIR RESULTS

Evaporation—Condensation—Humidity—Dew—Fog—Frost—Rain—Hail—Snow—Artificial Ice—Distillation—Sublimation—Diffusion—Osmosis—Dialysis

Evaporation compared with vaporization.—Evaporation, like vaporization, is a process in which liquids pass to a gaseous state, but the term vaporization is used when the formation of vapor takes place throughout the entire mass of the liquid and evaporation when it occurs only at the surface of the liquid. In order for vaporization to occur, a liquid must be heated to its boiling temperature, evaporation takes place to some extent at all temperatures, but is, of course, increased by a high temperature.

Why evaporation takes place.—As the molecules composing a liquid are in constant motion and their attraction for each other, which holds them together, is only fairly strong, some of the molecules at the surface escape from the liquid into the air. Many of these are drawn back again by the attraction of the molecules at the surface, but more molecules escape than fall back and thus after a time all the molecules of the liquid escape and move more quickly and fly apart and thus become gaseous.

Factors which influence the rate of evaporation.—Important factors in influencing the rate of evaporation are: (1) The nature of the liquid. To see this, expose in separate dishes a little water, alcohol, and ether. It will be seen that the liquids which have the lowest boiling point evaporate the quickest. (2) The temperature of the liquid. Heating a liquid, by increasing the rate of molecular motion, hastens evaporation. (3) The amount of surface exposed. Evaporation takes place from the surface of a liquid, consequently, the greater the extent of free surface, the more rapid the rate of evaporation. (4) Pressure. In order to pass into the air, the molecules from a liquid must overcome the pressure of the air upon the surface of the liquid, if therefore there is no air, as in a vacuum, evaporation will occur much more rapidly at a low temperature than under ordinary circumstances. For this reason, vaccuum pans are used in the preparation of condensed milk, much of the nutrient of which is lost if it is subjected to a high temperature. (5) The amount of vapor already in the air, *i. e.*, the degree of humidity under ordinary conditions of temperature and pressure, the atmosphere can hold only a certain amount of vapor; for this reason, evaporation will not take place as quickly on a damp, as on a fine, day; *e. g.*, wet clothes dry more slowly on a damp day. (6) The rapidity with which moist air is driven away from around the evaporating liquid—*e. g.*, on a windy day, clothes will dry more quickly than when there is no wind, and fanning a moist surface hastens its drying.

Result of evaporation on temperature.—A surface from which evaporation takes place and the air around it are cooled by evaporation. Thus the air is cooled after rain while the moisture is evaporating. The evaporation of sweat from the surface of the body is one of nature's principal means of preventing body temperature becoming too high, and bathing the body with water or alcohol, while exposed to the air, so that evaporation will take place rapidly is a means often employed for reducing body temperature in fever. It is, as shown under vaporization, due to evaporation that the temperature of boiling water will not exceed 212° F. and it is because of the evaporation that goes on in the lower part of a double boiler that food does not burn in the upper part.

Why evaporation causes cold.—One reason for this is that it is the hottest molecules that move the fastest and consequently it is the hottest ones that escape first into the air. Another cause is that mentioned under vaporization, viz., liquids must have heat in order to become vaporized and they take it from the most available source.

Liquids, such as ether and alcohol, which evaporate more rapidly than water create a greater degree of coolness than the latter during evaporation since the warmer molecules escape more rapidly, and, also, a definite amount of heat being required to produce evaporation, the more rapidly the work is accomplished, the more rapidly the heat will be taken. Certain substances, as ethyl chlorid, which evaporate very quickly will, if sprayed on a part of the body, cause it to freeze in a few seconds.

Evaporation of solids.—As a rule, solids become liquefied before being vaporized, but there are a few solids that will pass into a gaseous state and disappear without forming any visible vapor. For example, wet clothes that are hung out in winter and become frozen will lose the ice that formed upon them and become quite dry, though the temperature is too low for actual melting to have occurred. Also certain solids, such as camphor, will volatilize and disappear without apparently liquefying. A few other solids, such as arsenic and iodin, will, if heated, be converted at once into vapor.

Condensation

Nature and cause of condensation.—Condensation is the process by which gaseous matter is changed to liquid and the changes that occur are exactly the opposite of those which take place in vaporization; viz., due to the application of cold, molecular motion becomes less and less, the particles become packed closer together, and the heat of vaporization which was employed in keeping them apart passes from the matter.

A common use of the heat given off in condensation.—Heating houses by steam is a common use to which the heat given off in condensation is put. The steam generated in a boiler passes through the pipes with which the rooms are supplied and there condenses; thus the pipes, which radiate the heat they receive to the room, are heated, not only by contact with the hot steam, but also with the heat liberated during condensation. The water thus formed passes through pipes provided for the purpose, back to the boiler where it is once more converted into steam.

Liquefied gases.—Aqueous or water vapor is as easily changed to water as water is to vapor, but gases such as oxygen and hydrogen can be liquefied only at extremely low temperatures and under great pressure, and, as soon as the pressure is released, the liquid expands and the natural form of the gas is resumed. Such volatilizing matter takes the heat that it requires for the process from anything with which it comes in contact; for this reason, men, who have not taken proper precautions, have had their hands frozen while repairing tanks containing liquefied gases.

Artificial ice.—The use of ammonia in the making of ice is a common practical application of the alternate condensation of gas and vaporization of the liquid obtained by condensation. Ammonia gas is generally used for the purpose in preference to other gases because it is easily condensed by pressure at ordinary temperatures to a liquid which boils, and consequently, vaporizes at a low temperature. The principles of an ice-making machine and the process of making ice are as follows: There are two large tanks containing coils of pipes. The coils in the two tanks are connected by pipes to which a pumping machine is attached. In one tank are large cans which are filled with pure water and surrounded with brine; this is generally made of calcium chlorid and water. The coil in the other tank is surrounded with water. The ammonia gas, which has been liquefied by pressure, is pumped into the pipes in the tank of brine, where, the pressure being reduced, the liquid volatilizes and takes the heat required for the process from the brine, thereby reducing its temperature to about 18° F. (it will be remembered that brine can be reduced to a very much lower temperature than water without freezing). The intense cold of the brine gradually freezes the water in the cans into solid blocks of ice. The ammonia gas passes from the pipes in the brine tank and by means of the compressor or pump is forced into the pipes in the other tank at such a rate that a pressure of 155 pounds is obtained and this liquefies the ammonia, which is then once more passed into the coils in the brine tank where the pressure is less and it volatilizes and so on. The water in the condensing tank is for the absorption of the heat given off by the gas as it becomes condensed.

Refrigerators and cold-storage rooms are cooled in much the same way as ice is made, the essential differences being that the cans of water are omitted from the brine tank and the brine, after being reduced to a temperature of about 16° F. is pumped into pipes that extend from the tank to coils of pipes in the ceilings of the places to be chilled. The brine is kept in circulation between these pipes and the tank where it is cooled thus the pipes are kept constantly cold and they keep the place in which they are situated cold in just the same manner as ice would.

The air of theaters is sometimes cooled by forcing evaporating ammonia or carbon dioxid through pipes within the room. The principle is the same as in making ice, the essential difference being that the heat for the volatilization of the gas is taken from the room instead of from brine. The pipes in which the volatilization takes place may become as cold as 6° F.

Humidity

The term *humidity* is used in referring to the moisture of the atmosphere. The moisture of the atmosphere is the result of evaporation from bodies of water, moist earth, snow, or ice.

When there is as much moisture in the air as the latter can hold it is said to be saturated. How much moisture the air can hold without becoming saturated depends upon the temperature. Naturally, as warm air is expanded and its molecules are thus farther apart than in cold air, warm air will hold more moisture than cold air without becoming saturated. Cooling of the air, especially when it is saturated, is likely to result in condensation of the vapor and consequent forming of dew, fog, rain, etc., which, depending upon the temperature and where and how the condensation occurs.

Absolute and relative humidity.—By absolute humidity is meant the amount of aqueous vapor which the air contains expressed in the number of grams of moisture per cubic foot, or other specified amount, of air. By relative humidity is understood the amount of vapor that is present in the air expressed as a percentage of the amount that the air will hold without the moisture being precipitated in the formof dew, rain, etc.—*i. e.* without becoming saturated. When air is saturated, the relative humidity is said to be 100 per cent. How much moisture there can be in the air without its becoming saturated will, as stated in the preceding paragraph, depend upon the temperature.

The degree of humidity is ascertained by the use of an instrument called a *hygrometer*.

Effect of moisture on comfort and health.—The estimates given for the most desirable relative degree of humidity vary between 50 and 70 per cent. A higher per centage than seventy-five makes both cold and heat harder to bear, and it, when the temperature is high, increases the danger of heat prostration and sunstroke, because, for the reasons given under evaporation, it interferes with the evaporation of sweat from the body. Very dry air is also exceedingly injurious for it acts as a mental irritant and it parches the membranes of the eyes, nose, throat, and lungs, and thus lays a foundation for catarrh and bacterial infection. It has been found that steam-heated dwelling houses often have a relative humidity as low as thirty per cent. when the humidity out of doors is seventy-five per cent.

A common device to keep the atmosphere of artificially heated rooms moist is to have bulbs, or other water plants growing in wide shallow basins of water. The constant evaporation that occurs helps very considerably in keeping the air moist.

Nature of dew, frost, fog, clouds, rain, hail, snow.— When moisture in the air comes in contact with a surface cooler than the air, the moisture will condense upon it. The earth, grass, flowers, etc., are likely to become cooler than the air at night, for though they absorb heat during the day, they radiate it very quickly after the sun goes down, consequently the vapor in the atmosphere condenses upon them as drops of water. This is known as *dew*.[1] When the cooling that occurs is so great that not only the ground and plants, but also the air above them, become colder than the higher strata of air, the moisture in the colder portion condenses on all the particles of dust and the like suspended in the air and the result is known as fog. Naturally, fogs occur most frequently near large bodies of water and following days in which heat has caused a large amount of evaporation or where, as in parts of California, the hot air of neighboring deserts becomes so expanded that its pressure is reduced to such an extent that it, by suction, draws forward the

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[1] The collecting of drops of water on the outside of a pitcher containing ice water, when it is in a warm place, is the result of the same cause as dew.

moist air from the Pacific where the pressure is greater. When this air becomes at all chilled, the moisture condenses and envelops the land in mist or fog. Another factor influencing the amount of fog likely to occur in a place is, naturally, the amount of dust present, since it is upon the dust that the water condenses.

Clouds, like fog, consist of particles of dust upon which water vapor has condensed, but clouds are higher in the air.

When the condensation of water in the clouds reaches the point of saturation, the water falls upon the earth in the form of rain unless (1) the temperature of the air in which the clouds are, is below freezing point, when the water falls as snow, or (2) if the raindrops are whirled by the wind through regions of different temperatures some of which are below freezing point the drops of water fall as the small balls of ice known as *hail.*

Distillation

Nature of distillation.—If the vapor of a boiling liquid passes into a tube or pipe and is collected in a reservoir surrounded with ice as demonstrated in Fig. 40, the vapor will condense. This process is known as *distillation.*

Fractional distillation.—Since, as shown in experiment 6, liquids of different specific gravities have different boiling points, one with a lower boiling point than other substances in a compound can be extracted by keeping the latter at the boiling temperature of the liquid desired. For example, alcohol can be extracted from water by distillation, if the solution is kept at 80° C., the boiling point of alcohol.

The extraction of a liquid from a compound in this way is spoken of as *fractional distillation*.

Destructive distillation.—When certain solid or liquid substances, as coal, wood, and petroleum are heated in closed vessels out of contact with the air, they break down into simpler substances some of which can be easily vaporized, and, therefore, distilled. This is termed *destructive distillation*. The term *dry distillation* is sometimes used, for destructive distillation of solid substances.

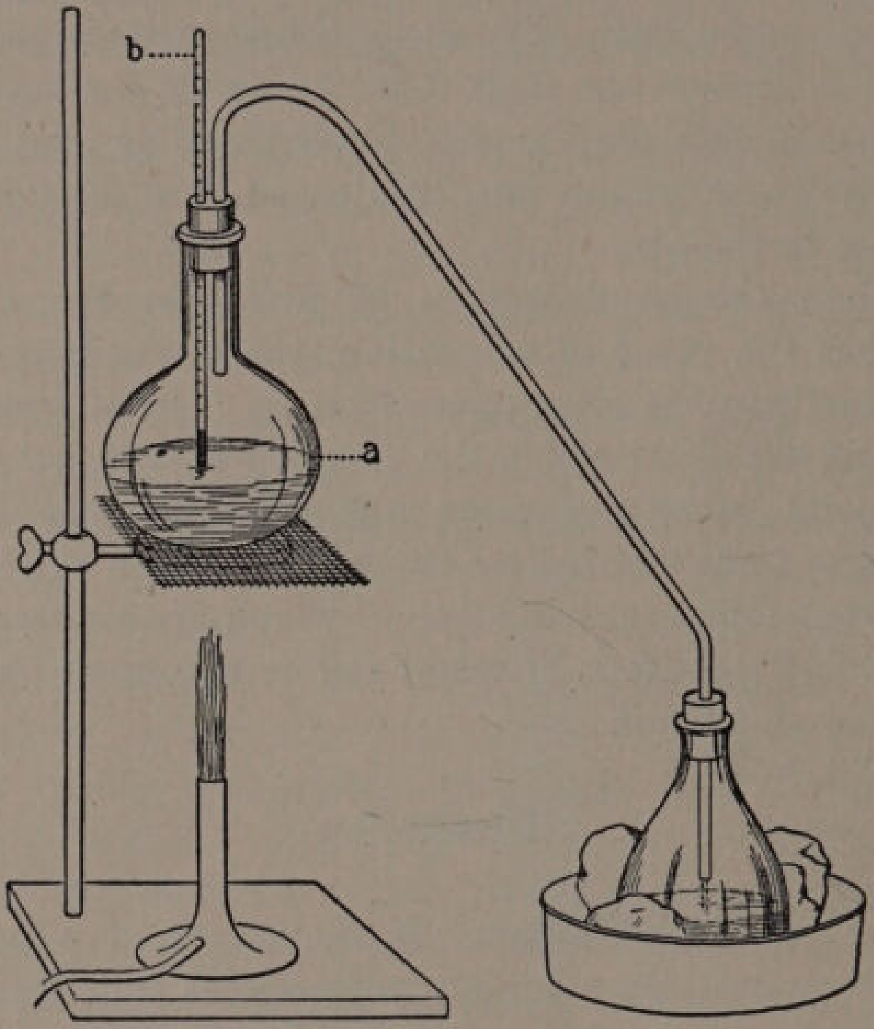

FIG. 40. DISTILLATION.
(a) Flask containing 50% Alcohol. (b) Thermometer.

Sublimation

Nature of sublimation.—In the section on the evaporation of solids it was stated that there were certain solids which passed into a gaseous state without any visible formation of liquid. If the vapor arising from these substances is collected and condensed it will solidify. This form of distillation is known as *sublimation* and the condensed products are called *sublimates.* Sometimes chemicals that will, when heated, interact to form new compounds are sublimed together. This is the manner in which the well-known disinfectant, bichlorid of mercury is prepared. The chemicals used in its preparation are mercuric sulphate ($HgSO_4$) and sodium chlorid ($NaCl$). When proper proportions of these substances are heated together, the mercury and chlorin unite in the proportion of one part mercury to two parts chlorin and sublime while the sulphate and the sodium unite to form a non-volatile substance—sodium sulphate—which remains behind. Thus $HgSO_4 + {}_2NaCl = Na_2SO_4 + HgCl_2$.

Diffusion

Nature of diffusion.—When two or more gases that have no chemical affinity for each other are introduced into a room, vessel, or other space, the gases quickly spread and intermingle so that in a short time each gas is uniformly distributed throughout the whole space; just as it would be were it the only gas present. This spreading of gases is known as the *diffusion of gases.*

Common proof of the diffusion of gases.—The rapidity with which odors from strongly scented matter become distributed is a very common proof of the diffusion of gases.

Experiment 7. Object, to observe the diffusion of gases.

Articles required: Two drinking glasses, the same shape and size, two swabs, a piece of stiff, smooth paper slightly larger than the open ends of the glasses, ammonia water (NH_4OH) and hydrochloric acid (HCl).

Procedure: Wet the interior of one glass with ammonia, using a swab for the purpose, and cover the glass with the paper, swab the inside of the other glass with hydrochloric acid, invert this glass over the first one and then remove the paper, being careful not to separate the glasses while doing so. It will be noticed that, shortly after the paper is removed, the glasses become filled with a white cloud. This is ammonia chlorid, derived by the combining of the ammonia gas from the ammonia water and the chlorin gas from the hydrochloric acid. Thus $NH_4OH+HCl=HN_4Cl+H_2O$.

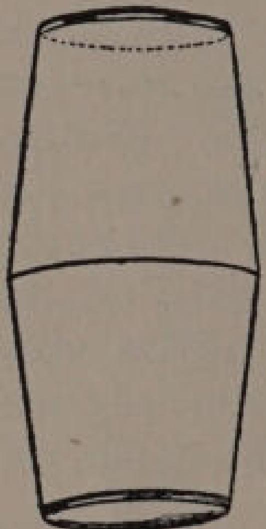

FIG. 41. GLASSES IN POSITION FOR DIFFUSION EXPERIMENT.

Diffusion of liquids.—If two liquids, for example lemon juice and water, are put into a glass they will, after a short time, become thoroughly mixed, even though they are not stirred. The diffusion of liquids, however, will not take place as quickly as did that of the gases, since the molecules of liquids are not in such rapid movement as those of gases, and diffusion is dependent upon molecular motion.

Diffusion of solids.—It has been found that if a mass of lead is placed upon a mass of gold, molecules of the latter will after a time be found scattered through the lead. The diffusion of solids, however, at ordinary temperatures is rare, but it is common at high temperatures.

The phenomenon of diffusion is considered one of the strongest proofs of molecular motion.

Osmosis

Nature of osmosis.—When matter diffuses through a membrane the process is spoken of as *osmosis*. Substances having a crystalline form, as sugar, will readily go into solution and osmose through membranes, but matter that is colloidal (as glue) or amorphous (*i. e.*, not crystalline, without definite shape) is not capable of osmosing. It is for this reason that animal foods, vegetables, and starches must be digested and thereby made soluble, before they can be absorbed.

Experiment 8. Object: To study osmosis.

Articles required: 2 gold-beater's bags or fish bladders, 2 corks with a hole though the center, 2 pieces of glass tubing of small caliber, about 9 inches long, concentrated sodium chlorid, water, two jars or wide-necked bottles.

Procedure: Fill a bag with salt solution, fit a cork, with a glass tube inserted in the hole through its center, into the open end of the bag, and tie it in place. Suspend the bag in a jar of

FIG. 42.
JAR WITH MEMBRANE BAG SECURED IN PLACE.

water, as in Fig. 42. Arrange a duplicate apparatus, but put the sodium chlorid into the jar and the water into the bag.

After some hours it will be noted that the solution has risen in the tube of apparatus 1, showing that water must have entered the bag, but that the solution has risen in the jar of apparatus 2, showing that water has passed from the bag into the jar.

Experiment 9. Object: Same as experiment 8.

Procedure: Have a carrot, instead of the membrane bag, and syrup, instead of salt solution. Scoop out the interior of the carrot, fill the cavity with a thin syrup; insert the cork with the tube inserted in the hole through its center at the top of the cavity letting the tube extend about an inch into the syrup. Suspend the carrot in a jar of water. Some of the syrup will pass into the water, but a much greater amount of water will osmose into the syrup; enough, in fact, to cause the syrup to rise in the tube.

Facts shown by experiments.—These experiments prove that substances in solution can pass through animal and vegetable membranes, and that a fluid of low specific gravity will osmose more readily than one of higher concentration. This fact is of particular interest to nurses who prepare salt solution for intravenous infusion; for if too much salt is used (*i. e.*, more than enough to make the solution stronger than 0.9 per cent.) fluid will be extracted from blood corpuscles and they will become shrunken and shriveled. If, on the other hand, the solution contains too little salt (less than 0.6 per cent.) it will enter blood corpuscles and cause them to swell and perhaps rupture. If a little defribrinated blood can be obtained from the slaughter house,[1] this effect can be easily demonstrated as follows: Put about 5 cc. of blood into each of four test tubes, to one of these tubes add an equal amount of water to another an equal amount of a 0.2 per cent. salt solution, to the third physiological salt solution (*i. e.*, 0.9 per cent.) and to the fourth, add the same quantity of 5 per cent. salt solution. Notice the difference in the color that occurs in all the tubes, except that with the physiological salt solution, due to the action of the salt solution on the blood corpuscles. The disintegration of the red corpuscles is spoken of as *hemolysis* and blood in which hemolysis has occurred is said to be *laked*.

For further study of the action of liquids of different concentration upon the blood, prepare a drop of the contents of each of the tubes for examination under the microscope. To do this smear the drop over the center of a clean glass slide. Place this so that it can be viewed with the microscope and note the difference in the shape of the blood-cells in the different smears.

The blood can be also prepared for examination as follows: Wash the finger with alcohol, prick it with a sterile needle, put a drop of blood thus obtained on each of four sterile slides. To one drop, add a few drops of water and to each of the others a few drops of one of the different percentages of salt solution. Mix the blood and diluent with a platinum tip.

A cause of edema.—Another fact of interest to nurses that the osmosis of fluid towards the location

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[1] When blood is stirred rapidly as soon as it is shed, the fibrin collects on the stirring implement. Blood thus treated is said to be *delibrinated* and it will not clot, for the fibrin, which has been removed, is essential for the clotting of blood.

of the more concentrated solution explains is that, when the kidneys fail to eliminate salt, as is often the case when they are diseased, and the salt passes into the tissues, œdema results, because of the consequent osmosis of abnormally large amounts of the fluid contents of the blood into the tissues.

Dialysis

Nature and use of dialysis.—Since all substances cannot pass through animal membranes, chemists often make use of what is known as a *dialyzer* (a jar with a parchment bottom) to separate non-crystalline from crystallizable substances when the former are present in a solution in a finely divided form. The process is known as *dialysis*.

CHAPTER V

SOME COMMON PHYSICAL PROCESSES AND THEIR RESULTS —(*Continued*)

Contraction—Expansion—Vacuum—Suction—Siphonage—Capillarity—Conduction—Convection—Some Common Practical Applications of Knowledge of the Nature and Action of these Processes.

Contraction and Expansion

It is a well-known fact that even solid substances are expanded by heat and contracted by cold. For instance, railroad rails are laid with spaces between their ends to allow for the expansion that occurs during the heat of summer; it is difficult when the hands are hot to put on kid gloves that can be drawn on easily when the hands are cold. Liquids cannot be expanded to as great an extent as gases since high degrees of heat and cold change their state, *i. e.*, change them to vapor or solid—but that some change is made can be easily seen if a graduated flask containing water is held over a flame, for the water will be seen to rise in the flask. The degree to which gases are expanded by heat and contracted by cold is stated in the law of Charles—$\frac{1}{273}$ of their volume for every degree centigrade that their temperature changes. The expansion of gases can produce great force as is demonstrated by explosions. For example, the dynamite used to blast out rock, etc., does so because a small amount of solid dynamite is converted by heat into large quantities of gas which, by the heat produced as the result of the chemical interaction of the constituents of the dynamite, becomes very greatly expanded and forces apart even huge rocks that interfere with its expansion.

Reason for expansion and contraction.—The change in the size or volume of matter as the result of heating is due to increase in the velocity of molecular motion by the heat in consequence of which the molecules jostle each other farther apart, thus increasing the volume of the matter. Cold, on the contrary, lessens molecular motion and, consequently, contracts the majority of substances.

Peculiarity of expansion and contraction in water.—Fortunately for mankind, water, when cooled, does not act in quite the same way as most substances; it contracts until a temperature of 4°C. is reached; but after it attains this temperature, it begins to expand, and consequently its weight grows less in proportion to its bulk. For this reason, when ice forms in lakes, rivers, etc., it floats to the top. If the ice did not do this, bodies of water would freeze to the bottom and, if deep, could not be thawed, even in summer.

Why glass and porcelain-ware break when heated or cooled quickly.—If glass and porcelain utensils are heated and cooled slowly they will stand wide variations in temperature without breaking, but, on the contrary, if they, especially glass ones, are heated or cooled quickly, particularly if the heat is directed against one point or if, when hot, they are placed upon a cold stand, they are likely to break. This is because, in such cases, the expansion or contraction is uneven and a crack occurs where resistance is offered to the change in condition.

If when pouring a hot liquid into a glass utensil, the liquid is poured over, as well as into it, there is less danger of its breaking and thin glass is broken less readily by heat than thick glass. The reason in both instances being that all parts of the utensil are brought more nearly to the same temperature at once.

The thermometer.—The thermometer is a practical application of the expansion and contraction of liquid by heat and cold. A thermometer usually consists of a glass tube of capillary bore with a bulb that is filled with mercury at one end. When making a thermometer, the bulb and part of the tube are filled with mercury and the instrument heated until the mercury boils and expels the air from the tube. After this the top of the tube is sealed. When the tube is cold, the bulb is placed in boiling water and the point to which the mercury rises is marked 212, if the thermometer is marked with the Fahrenheit scale, or 100 for the centigrade scale. The bulb is next placed in a vessel of melting ice and the point at which the mercury stops is marked 32 for the Fahrenheit scale and 0 for the centigrade. The space between the freezing and boiling points is divided into a number of equal spaces, the size depending upon the kind of scale.

Vacuum, Suction, Siphonage

Vacuum.—By a vacuum is meant *a space from which all air has been exhausted and which contains no material substance.*

A vacuum is usually created by means of some form of pump, fan, or piston which will drive air or other matter from a cavity.

Suction.—By suction is meant the drawing of matter into a vacuum. Suction depends upon three things, viz.: (1) the tendency of gases and liquids to diffuse equally in all directions; (2) the fact that, when a vacuum has been created within a space or cavity, there is no pressure within it to oppose the entrance of any form of matter; (3) the pressure of the atmosphere upon the matter that is afforded entrance.

Some common uses of suction.—Suction, due to the creation of a vacuum, is an essential principle in the raising of water from a lower to a higher level and in drawing fluids from and into cavities, it is the power that is being more and more used for the removal of dirt and dust from buildings and streets. Its chief values for the latter purpose is that the dust, etc., are removed without being scattered or inhaled and, in buildings, carpets and the like are not subjected to the same amount of wear that they are by other modes of cleaning. Suction is also one of the forces depended upon in the ventilation and heating of buildings as will be seen in the section on ventilation.

Siphonage.—When a liquid is drawn through a tube over an elevation to a lower level, the process is spoken of as *siphonage* and the apparatus used for the purpose as a *siphon.*

Siphonage forms a part of many of the procedures used in the treatment of disease. The necessary vacuum in the tube through which the fluid is to flow is generally created either with a syringe, or small pump, or by letting liquid run through tubing and thus forcing out the air. There are two things necessary in order to get good siphonage by this last method, viz.: (1) After a vacuum has been created in the tubing the entrance of air must be prevented; (2) the end of the tubing from which the fluid is to flow after it has been extracted from the reservoir or cavity must be considerably lower than the end that is in the reservoir or cavity. A good way to start siphonage from a pail or similar reservoir, when there is no syringe at hand, is to insert a funnel in one end of a piece of tubing and, holding the tubing so that its free end is on a level with, or a little above, the funnel, as in Fig. 43, to pour water, or whatever liquid is being used, from a pitcher into the funnel until it appears at the opening of the free end of the tubing and then to, *simultaneously*, lower the funnel into the pail of liquid and the free end of the tubing over the vessel or part into or over which the liquid is to run. This must, as previously stated, be lower than the pail. In doing this an important precaution is to prevent the tubing being bent upon the rim of the pail.

FIG. 43. MANNER OF HOLDING FUNNEL AND TUBING WHILE FEELING THEM.

Capillarity

Nature of capillarity and reason for name.—If a tube of small diameter is placed in a liquid that will wet its interior, the liquid will rise somewhat higher in the tube than the level of the liquid on the outside, and the smaller the diameter of the tube, the higher the liquid will become. If the tube is placed in mercury or other liquid that will not wet it, the liquid will become depressed, rather than elevated. The force causing a liquid to rise in a tube is known as *capillarity* or *capillary attraction* (from the Latin *capillaris* = hair-like) because it is only perceptible in tubes of small diameter.

Cause of capillarity.—The molecules of a liquid have an attraction for each other which holds them together (cohesion) and also an attraction for the molecules in the sides of a tube which they wet (adhesion—see page 36). If the force of cohesion is stronger than that of adhesion the molecules of the liquid will be pulled down, but when adhesion is the stronger force the liquid clings to the sides of the glass and rises in the tube.

It is due to capillarity that oil rises in the lampwick; that a piece of muslin or other tubular fibered fabric will become wet throughout if a small end is left in a liquid; that gauze in a wound will act as a drain. Capillarity plays an important part in the flow of sap through plants and in the distribution of moisture through the soil.

Conduction

It was stated in a preceding chapter that heat was transferred from one body or part to another by either one of the three processes conduction, convection, or radiation.

The term conduction is used also in connection with the transmission of electricity, nerve impulses, and sound.

Nature of conduction.—Conduction as applied to heat, has been defined as *the transfer of molecular motion from a mass of molecules vibrating with a given intensity to another mass vibrating with less intensity.*

Experiment 10. Object, to study the method of heat conduction.

Procedure: (a) Hold one end of a piece of metal wire, preferably copper, in a flame. How long can you hold it thus before the end between your fingers becomes too hot to hold?

(b) Hold a teaspoon by the handle and keep its bowl in boiling water. How long does it take for the handle to become too hot to hold?

(c) Fill, to $\frac{3}{4}$ its capacity, a large-sized test-tube with water, hold the tube, at the lower end, in the fingers and heat the water in the upper part by holding it in a flame. Notice how much longer it takes for the water in the bottom of the test-tube and even for the glass of the tube to become hot than it did for the wire or the spoon.

(d) Refill the tube with cold water and, holding it by the upper end keep the lower part in the flame. Notice what a very much shorter time it takes for the water to be heated in this way, which is convection, and not conduction. This will be discussed in the following section.

Explanation of experiment.—The heat of the flame caused the molecules of matter within it to vibrate violently and these molecules, by striking against those in the part of the spoon, wire, etc., held in the flame, furthered the increase of molecular vibration that had already been increased by the heat. The vibration of the molecules in this part of the metal started those above them vibrating and these set the masses above them vibrating and so on.

As seen by the experiment, metal is a better conductor of heat than glass or water. This is because the molecules are packed more closely together in metal than they are in the other two substances used and thus they can influence each other more readily. Metals are the best heat conductors, but all solids are better than liquids and liquids than gases. The following table gives the relative heat-conducting value of a few common substances. The one having the highest value is, for convenience, marked 100.

Silver	100	Mercury	1.7
Copper	74	Granite	0.53
Aluminium	35	Limestone	0.52
Brass	27	Ice	0.21
Zinc	26	Glass	0.2
Tin	15	Water	0.124
Iron	12	Felt	0.038
German silver	8.4	Air	0.005

A few of the common applications of a knowledge of the relative heat-conducting power of matter to household purposes.—A common application of this knowledge is seen in refrigerators, ice-houses, and fireless cookers. The walls and doors or covers of these appliances are made with double walls that have a space between them. The space is filled with air or other non-conducting matter. An improvised fireless cooker is sometimes made by taking two boxes, one considerably smaller than the other, and filling the space between the two with some such porous, poor heat-conducting material, as sawdust, wool, or shredded paper. It must be possible to close the covers of the two boxes securely and a cushion, stuffed with the same kind of material as that between the boxes, is placed between the covers. If ice—in a bowl or basin—is placed in the inner box, the latter can be used as a refrigerator or if, instead of ice, a covered pot of boiling food is placed in it, the food will continue to cook.

The walls of all such appliances must be air-tight for they will be valueless if air is allowed to enter or leave the space, since this would start convection currents, and though air is one of the poorest heat conductors known, it distributes heat very readily by convection. If, however, air currents are prevented, the temperature of the refrigerator or cooker will change very slowly.

The thermos bottle, which consists of two bottles blown one inside the other, is made on the same principle as the refrigerator, but the space between the bottles is a vacuum, the air being pumped out before the bottles are sealed together at the top. Naturally, heat cannot be conducted through a vacuum and thus it is only around the cover that heat can pass or enter a thermos bottle and thus hot liquids put into them will remain hot, and cold ones cold, for several hours.

Flannel and other loosely woven woolen material owe their value for warm clothing largely to the air which is held within the meshes of the material, this interfering with the escape of heat from the body.

Convection

Definition.—Convection has been defined as *the transmission of heat by the transfer of the heated body itself.*

Nature.—In experiment 10, it was seen that if the upper part of a test-tube filled with water was held in the flame, the water at the bottom of the tube did not become warm for a long time, but that if the bottom of the tube was held in the flame, all the water soon became heated. The reason for the first fact was explained under conduction. The reason for the second is that as heat expands matter and consequently makes it lighter in proportion to its bulk than that which is cold, the heated water in the bottom of the tube became lighter than that at the top and, just as a cork, being lighter than water, will, if thrust to the bottom of a basin of water, rise and float, the lighter water rises above that which is colder and consequently denser and heavier. This results in the pressing down of some of the colder water toward the source of heat. Some of that pressed down at once becomes warmer than that directly above it and rises with the same results as before. Thus, there is a constant movement in the water and in time it all becomes heated by, as stated in the definition, *the transfer of the heated matter itself*, or, in other words, by convection. This transmission of heated matter occurs not only in water, but in all fluids, both liquids and gases.

Convection currents in nature.—Winds are convection currents produced by differences in the density and pressure of the air. Such differences are usually the result of differences in temperature due to unequal heating of the earth and air by the sun. When air in any locality is heated it rises and, the atmospheric pressure near the earth being reduced, air from colder regions passes in. This movement of the air constitutes a breeze or, if there is enough difference in the temperature to promote rapid motion, a wind. The land bordering on large bodies of water will, usually, during the day become hotter than the water, consequently the air over the land becomes warmer than that over the water and rises and the colder, denser air from the water moves forward over the land. This constitutes what is known as a *sea breeze.* If, as usually occurs at night, the air over the land cools before that over the water, the air passes from the land toward the water thus giving rise to the so-called *land breeze.* The trade winds, that make life in the tropics endurable during the hot seasons, are the result of the movements of the hot air, which, rising, passes to the north and south leaving space for the entrance of cooler air from the latter places.

The various ocean currents that are of such importance to the climate of many regions of the globe are convection currents due largely to different densities of the water, resulting from its unequal cooling and heating in different parts, and to the rotation of the earth.

Ventilation.—The two chief forces depended upon in the ventilation of buildings are convection and suction.

A good illustration of combined convection and suction is seen in the open fireplace. The fire heats the air in and around the fireplace thereby making it lighter and causing it to rise through and from the chimney. This creates a partial vacuum in the chimney and fireplace which the colder and, consequently heavier air from outside and from the room is drawn in to fill. The same forces are at work also when radiator pipes are the source of heat and open windows the means of ventilation. The heat radiated from the pipes causes convection currents in which the heated air rises and is forced, so great is the pressure of heated gases, even through cracks around windows and doors, thus the air pressure in the room is reduced and the colder air is sucked in.

The locations where the greater amount of hot air will naturally escape and cold air enter can be seen by, on a cold day, opening a window at the top and at the bottom and holding a candle in front of first one and then the other opening. It will be seen that, at the lower opening, the flame is forced inward by the incoming air and, at the upper opening, it is forced outward by the outgoing air.

Knowing these facts, it will be easily appreciated, that, if it is too cold outside to open a window wide, it is better to open it slightly at both top and bottom. A very important reason for an opening at the top of the window is that the air exhaled in respiration, being warmer than the air in the room, and gases that may escape combustion when gas is burned tend to rise.

The openings for the entrance and exit of air should never be placed directly opposite each other, for then two things will occur, (1) a wind or draught may be created and (2) the air passes directly from one opening to the other and thus that in parts of the room not in the direct current is little affected.

Artificial ventilation.—In cold climates, in order to secure good ventilation in large buildings without risk of draughts from open windows, different forms of artificial ventilation are now in common use. The various methods by which ventilation is thus affected are classed under three headings, (1) extraction or vacuum system, (2) propulsion or plenum system,

(3) vacuum-plenum system. Ventilation by any of these methods is usually effected with the aid of fans or pumps placed in chimneys or ventilating shafts. In the vacuum system, air is withdrawn from the building as the result of the vacuum created in the shafts and the pipes leading to the rooms, by means of the pumps, etc., and air from outside is consequently sucked through the pipes and ventilators provided for the purpose. In the plenum system, air from outside, that is drawn into the ventilating shaft, usually by the creation of a vacuum, is forced through pipes or shafts into the rooms, the entrance of this additional air forces some of that already in the rooms to leave through the apertures provided for the purpose. The vacuum-plenum system, as the name implies, is a combination of the other two methods.

One means of discovering if artificial ventilation is proceeding as it should is to hold a handkerchief in front of the ventilators; if everything is in order, the handkerchief will be sucked toward a ventilator through which air should leave the room and it will be blown into the room by the incoming air at a ventilator intended for such purpose.

CHAPTER VI

THE ETHER, HEAT, AND LIGHT

The Ether—Absorption, Radiation, Reflection, Refraction, and Polarization of Heat and Light—Color of Light and of Objects—Finsen Light—Phosphorescence and Fluorescence.

Radiation

Nature of radiation.—The process of radiation was so named because heat and light transmitted in this way pass from their source in straight lines or *radii*. One proof that this is true of heat is demonstrated by holding a screen between the individual and the source of heat, since the direct heat then ceases to be felt. That any heat is then perceived is due to convection currents in the air. Likewise, if an object through which light rays cannot pass is held between the eyes and the light, the latter can no longer be seen and that there is any lighting of a locality screened from the direct rays of a light, is due, as will be seen later, to reflection.

The Ether

In order to understand radiation it is necessary to know of the existence of the ether. For various reasons, among others, the transmission of heat and light through the millions of miles between the sun and the upper strata of air and the nature of radiation, scientists concluded that there must be something besides air permeating all space, not only that above the air but also that existing in all matter, and to this invisible, intangible something they gave the name of *ether*.

Nature of ether waves.—All matter, but more especially that in which molecular motion is pronounced enough to cause sensible heat, will, by means of its molecular motion, start waves in the ether that, though not precisely similar, are likened to the series of circular waves that arise when a stone is cast upon the quiet surface of a pond. These waves spread out in all directions and, the hotter the body starting the vibrations, the more rapidly will the waves be formed and the greater will be the number of short waves (*i. e.*, those in which there is only a short distance between successive waves).

Length of ether waves.—Ether waves started by very hot bodies, such as the sun, vary greatly in length. Some, for example those made use of in wireless telegraphy, are miles long and others are only about $\frac{1}{250,000}$ part of an inch. As the longer waves produce electric effects, they are termed *electric waves*. Waves which measure only a few thousandth part of an inch are capable of increasing molecular motion in matter and consequently of raising their temperature; they, therefore, are called *heat waves*. Waves that are as short as $\frac{1}{30,000}$ part of an inch affect the endings of the optic nerve in the eye and are therefore called *light waves*.

Origin of color.—Light waves have different lengths and difference in length results in different colors.

Waves that are less than $\frac{1}{60,000}$ part of an inch do not affect the sight, but they do affect the chemicals upon a photographic plate and they cause many chemical changes in nature, therefore, they are called *chemical* or *actinic waves* or *rays*.

TABLE SHOWING THE WAVE LENGTH OF DIFFERENT COLORS OF LIGHT:

Color	Wave length[1]	Color	Wave length
Red	.0007 mm.	Green	.00053 mm.
Orange	.0006 mm.	Blue	.00047 mm.
Yellow	.00058 mm.	Violet	.0004 mm.

Only the molecules of very hot bodies vibrate with sufficient rapidity to affect the optic nerve (*i.e.*, to make light waves or rays). The temperature of a body must be raised to 525° C. before waves short enough to be visible as light appear. When this temperature is reached, a red glow is seen and as the temperature is increased, though long waves are being generated, shorter and shorter ones (those that give rise to the different colors) are produced until, between a temperature of 860° C. and 1200° C., waves of all colors are being generated and consequently a so-called *white heat* or light is reached for, when all colors of the spectrum[2] are present in a combined state, a white light is produced. This is the case with sunlight. The sun, being the hottest of all bodies, produces waves in the ether of all possible lengths; thus, in the sunlight, as already stated, there are waves that will give rise to electric effects, heat,

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[1] A millimeter (mm.) equals 0.3937 of an inch.

[2] The colors seen when light waves are refracted by passing through a prism, etc.

light of all colors, and chemical reactions. That the sunlight contains all the colors of the spectrum is seen when it undergoes refraction, as in the rainbow and in passing through prisms, see page 100.

In one sense, all ether waves are heat waves, since waves of all lengths will increase molecular motion in bodies and thus raise their temperature.

Retention of heat.—Light waves can penetrate many substances easily that the longer heat waves pass through only with difficulty and this is of great value to man. For instance, light waves pass through the glass of hot-beds and conservatories and through snow, paper, etc., covering plants, they are partly absorbed by the earth and plants, their molecular motion is thereby increased and heat thus produced which, as it does not readily pass through the plant covers, remains and furthers the growth of the plants. Also, the light waves pass readily through damp air, smoke, and the like, but the heat waves only do so slowly and with difficulty; for this reason, clouds and smoke at night serve to maintain the warmth that the earth acquired during the daytime. Thus a clear night is likely to be cooler than a cloudy one. The fruit growers of California burn smudge fires in their orchards when there is danger of a frost, because heat and also a heavy low-lying smoke that prevents the escape of heat from the ground results from these fires.

How the air is heated.—Heat being the result of molecular motion, it cannot, it is believed, exist apart from matter and consequently ether waves are not themselves hot, though, by increasing molecular motion in all matter which they strike, they are the cause of heat.

Ether waves would pass through absolutely dry air without producing any decided heating effect, but moisture and solid matter, as dust, in the air will give rise to heat. Moisture also serves to hold the heat and to transmit it in all directions. Also, as stated in the preceding paragraph, it helps the earth to retain the heat it acquires during the daytime and yet, it acts as a screen between the earth and the sun's rays. On high mountains, above the clouds, where there is little moisture, people will suffer intensely from the heat formed in their own bodies by the impinging upon them of the ether waves, while the air surrounding them is cold. These effects of moisture are one of the several reasons why the temperature of places near large bodies of water are more uniform the year round than those not so situated.

Effect of different kinds of matter upon ether waves.—When ether waves strike any medium other than ether, any one of three things may happen; they may be wholly or in part (1) reflected; (2) transmitted (*i.e.*, allowed to pass through the substance); (3) absorbed.

Opaque, translucent, and transparent objects.—Such substances as metals, wood, etc., which do not permit the passage of light rays are said to be *opaque*. Those which allow some light to pass through their substance, *e.g.*, ground glass and paper, but do not permit objects on their farther side to be seen, are said to be *translucent*. Objects, such as clear glass, through which light rays pass freely and allow objects on their farther side to be seen are said to be *transparent*.

Reflection

Nature of reflection.—Ether waves may be reflected or thrown back from matter which they strike in two ways; one of these is termed *regular reflection* and the other *diffuse reflection.* The straight beam of light that is seen striking back from a mirror upon which a ray of light falls on entering through a small aperture into a darkened room is an example of regular reflection. Diffuse reflection is the scattered, diffuse or irregular throwing back of light that occurs when light waves fall upon a rough or unpolished surface. It is due to the diffuse reflection of light waves from dust and moisture in the air and from objects upon the earth that there can be light all about us and not only in the direct path of the light rays. The phenomenon, popularly known as "the sun drawing water" is produced by the reflection of light, from sunbeams coming from behind a cloud, by dust and moisture in the air. The light of dawn and twilight are due to reflection of the sunlight by moisture and dust in the air which catch the light even before the sun appears or after it disappears below the horizon. The colors of sunset also are due to the reflection of light waves by the clouds and particles in the air. The colors are seen in the evening because the atmosphere is then denser than earlier in the day and thus more light waves are absorbed and, as the light waves then strike the clouds from below, their reflection is facilitated.

Why we see objects.—We see objects because they reflect light waves which fall upon them back into our eyes so that they strike upon the retina and form a picture there, as upon the sensitive plate of a camera, and thus stimulate the optic nerves which then transmit their impressions to the center of sight in the brain.

Differences in the power of absorption, reflection,

etc.—Matter varies greatly in the degree of absorption, radiation, and reflection of which it is capable and it is useful to have some knowledge of these differences. For instance, one without such knowledge might think that asbestos would be a good protector to put behind a stove, but asbestos, though itself non-combustible, absorbs and thus has its molecular motion increased and is thereby heated by ether waves and therefore it would be likely to scorch paper or wood behind it. A highly polished tin or zinc mat, on the other hand, reflects so many ether waves, instead of absorbing and being heated through by their action, that it will protect a wall behind it. White paint reflects more heat and light waves than dark paints, which is one reason why white or light colored paints, papers, etc., make a room lighter than dark colors and why black and dark colored clothing is warmer than white and light colors. Rough, dull surfaces absorb and radiate heat better than smooth polished surfaces and thus radiator pipes are always rough, to have smooth ones would entail a great loss of heat. Bright smooth surfaces, as mirrors and polished metal, are the best reflectors, but practically all bodies reflect some light waves, were this not the case, we would not see them, for, as stated in a previous paragraph, we see objects because of the light which they reflect into our eyes.

It is trying to the eyes to read what is printed or written on glazed paper and highly polished blackboards, because so much light is reflected from the shining surface that the words are obscured. An intense reflection or glaring light from any source is harmful to the eyes and the new method of artificial lighting, in which the light is thrown upon the ceiling, in order that it may be diffused and a more even, less concentrated, lighting be accomplished, is an attempt to obviate such defects.

The differences in the power of absorption, reflection, etc., are thought to be due, at least in part, to differences in the molecular motion of matter.

Origin of the color of bodies.—The color of objects depends upon the ether waves which they absorb and reflect; *e.g.*, grass is not green, but it appears to be green because it reflects green waves and absorbs those of other colors; a paint or dye appears to be a certain color because it reflects the wave of the particular color it appears to be and absorbs all other waves. Matter that does not absorb one set of wave lengths more than another, but reflects them all equally, appears white, and matter that absorbs practically all the ether waves incident upon it appears black.

Number of colors.—Though sunlight, when refracted by passage through a prism, appears to be separated into but seven distinct colors, many hundreds of colors are possible. One reason for this is that waves of more than one length are reflected by the majority of the chemicals used for dyes and paints and other objects and if waves of different lengths are reflected, but not in sufficient number or quantity to cause white light, a different shade or hue of the predominating color—*i.e.*, the wave length reflected to the greatest extent—can be obtained. As will be seen later, mixing paints alters the degree of absorption and reflection of waves and thus gives rise to different colors and shades.

Reason for changes of color under artificial lights.—Artificial lights, not being as hot as sunlight, do not produce all the wave lengths that the latter does and different kinds of artificial lights vary in the lengths of waves that they produce. Consequently, as objects have not the color, but essentials for color, they may not be the same color by artificial light as they are by sunlight, or by electric light as by gas light, for the particular wave or waves may not be present to be reflected.

Reason for the fading of colors.—The chemical rays of the sun and certain chemicals used for bleaching cause changes in the nature of the pigments used for dyes and paints which lessen their power of absorption of light waves, and thus more waves are reflected and this gives rise to the presence of a greater amount of white light, in consequence of which the color is lighter. The object is then said to be *faded* or *bleached*.

Why color is perceived.—In order that color be perceived by the brain it is necessary for the retina to be stimulated in some way by the color rays entering the eye. How it is stimulated is not known, but one theory advanced is that the retina contains a compound chemical substance which is acted upon by the incoming color waves in some way, as yet unknown, so that chemical changes are brought about which produce the sensations that give rise to the color sensations which we perceive.

Color blindness.—About 4 to 5 per cent. of men and 1 per cent. of women have defective color vision. As a rule, this defect consists in not being able to distinguish some one or, more frequently, two colors, but occasionally there is an inability to distinguish any color and individuals with this disability can distinguish form, light, and shadows, only. Any defect of color vision is spoken of as *color blindness.* The cause of the condition is unknown. Some authorities consider it due to a lack of some chemical substance in the retina and others to defective formation of the nerve-endings that should be affected by color.

Complementary colors.—When certain color waves, as red and green, yellow and blue, green and violet stimulate the retina at the same time, they produce a sensation of white. The pairs of colors which do this are called complementary colors. It is because yellow and blue waves produce a sensation of white that we use bluing to improve the *color* of white clothing that has become yellowed.

Reason for colors produced by mixing different colored pigments.—The results obtained when colored lights fall upon the retina are not the same, it will be observed, as when paints are mixed; for instance, when yellow and blue paint are combined we get a green color, because yellow pigment absorbs all colors except yellow and some green while blue absorbs all colors except blue and some green, and mixing the two pigments causes the absorption of all color waves except green. A white paint added to another lessens its power of absorption, hence more waves are reflected and a lighter color results. On the other hand, the addition of a so-called black pigment to a colored paint increases its power of absorption and the color appears darker.

Color of transparent objects.—The color of transparent objects depends upon the light waves they transmit—*i. e.*, allow to pass through them. All light waves pass through white glass, but if a coloring matter that will allow the passage of red waves only is added to glass in the process of its manufacture, then only red waves will pass through. Blue glass absorbs all color but blue waves and a little green and it transmits the blue and green. Yellow glass transmits yellow waves freely, green and red to a slight extent, but blue and violet not at all.

The power of colored glass to absorb some waves and transmit others is made use of in heliotherapy. By looking at the table on page 86, it will be seen that the waves which give rise to a red color are nearest in length to those that have the greatest power to produce heat in matter and, therefore, if red glass is placed between a patient and the sun's rays the waves passing through will have a heating effect, but those which produce chemical reactions and, consequently, so-called *sunburn* are shut out. Violet rays, being nearest chemical waves in length, the waves that pass through violet glass produce chemical reactions but have not such a heating effect as those which pass through red glass. The waves that pass through yellow and green glass give rise to slight heating and slight chemical effects. Thus yellow or green glass placed between a patient and the sun's rays will prevent sunburn and intense heating.

Finsen Light

When light is passed through a lens made of quartz or certain other substances, all ether waves are absorbed except those which give rise to violet and ultraviolet rays. The latter, partly, it is thought, by the chemical reactions that they promote in the tissues, are often very effectual in the treatment of such diseases as lupus. Ultra-violet light is obtained also by electrical means in somewhat the same manner as the X-rays, which will be discussed later.

The ultra-violet light was named *Finsen light* in honor of Niels Ryberg Finsen, the Danish physician, who first advocated their use in the treatment of disease.

Refraction

Nature of Refraction.—Ether waves of all lengths can be refracted, that is, bent from their course, in the same manner, but as it is the light waves that give the most obvious results of their refraction they are the ones that we will consider. The refraction of light waves occurs when they pass obliquely from a medium of one density into that of another density. If the medium into which the waves pass is of greater density than that in which they are—*e.g.*, when they pass from air into glass or water—they are bent toward the perpendicular, but when the medium into which they pass is of lesser density than that which

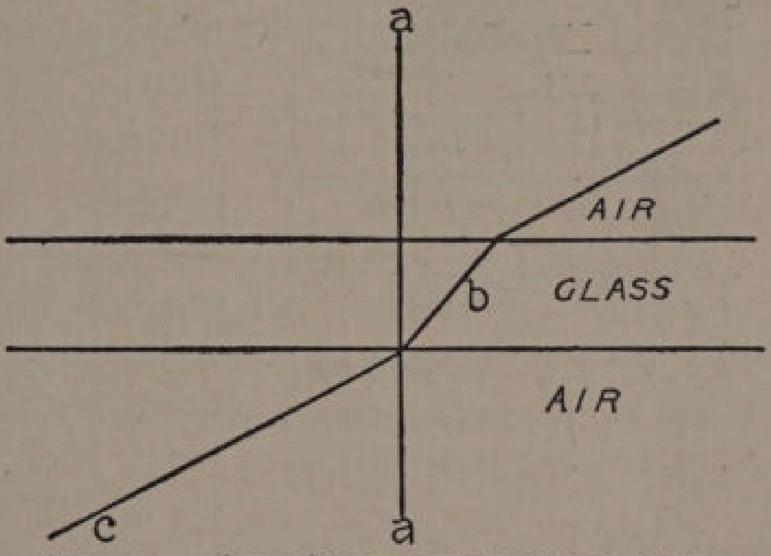

Fig. 44. Line Bent toward and away from the Perpendicular.

(a a) Perpendicular line. (b) Line bending toward the perpendicular. (c) Line bending away from the perpendicular.

they leave—*e.g.*, when they pass from glass or water into air—they are bent from the perpendicular. It is because the light rays coming from an object placed under the lens of the objective of a microscope are diverged, as in (b) Fig. 46, and thus enter the eye at a wider angle than the direct rays would, that the image of the object is magnified.

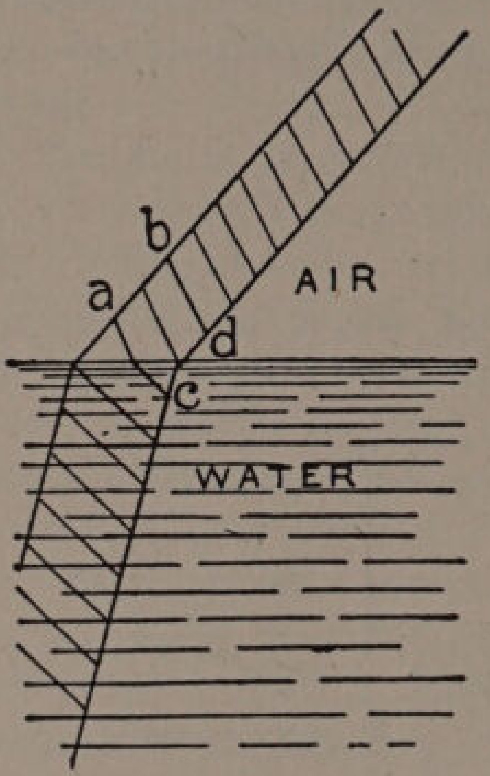

Fig. 45. The Bending of a Beam of Light Entering from the Air into Water.

Cause of refraction.—The cause of refraction is that ether waves travel at different rates in media of different densities; for instance, light waves pass through water at only three-fourths the speed with which they travel through air, and when a ray of light strikes a body of water obliquely all portions of its front will not enter the water at the same time and the part which enters first becomes retarded before that entering later. In Fig. 45 the short parallel lines represent the wave front of a beam of light entering obliquely from air into water; one part (a) reaches the water before (b) and, as it then travels more slowly than that part of the wave which is still in the air, it only reaches (c) while (b) is going to (d); thus the beam is bent. It is because of refraction that a stick held partly in the air and partly in water looks as though it were bent. Rays of light entering from one medium into another perpendicularly are not bent. It is due to refraction that the rays of light which are reflected by objects into our eyes both perpendicularly and obliquely come to a focus on the retina. The cornea, lens, and humors of the eye serving to cause refraction.

When the surface of a refracting medium is convex —*i.e.*, thicker in the center than at the edges—light rays come to a focus quicker than when it is flat, and the normal eye is provided with a means of becoming more convex when looking at objects near at hand.[1] When, for any reason, the eyes lose their

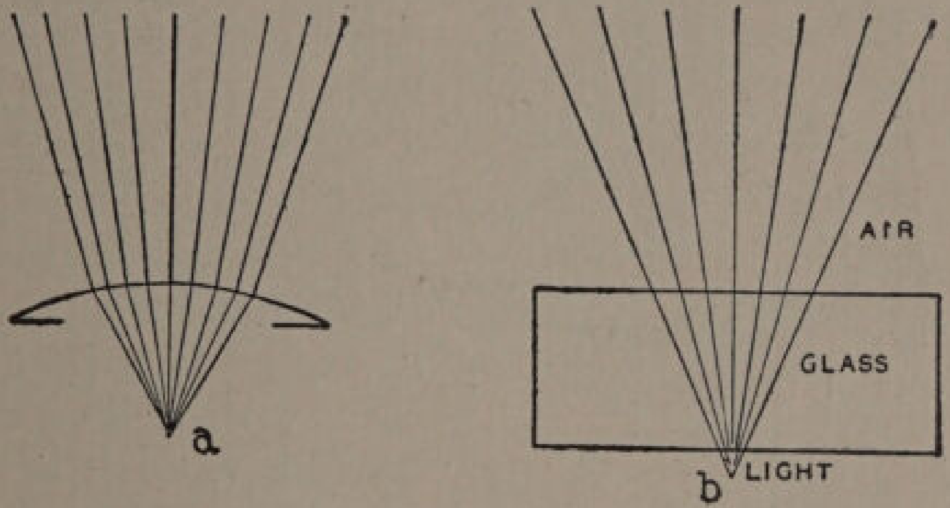

FIG. 46. (a) Showing how refraction by a convex surface will bring light rays to a focus. (b) Showing how light rays will be diverged on passing from a medium of greater into one of lesser density.

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[1] It is thought, that when the eye is at rest or fixed upon distant objects the suspensory ligament (*i.e.*, the ligament that is fused to the anterior surface of the capsule inclosing the crystalline lens of the eye and attached at either end to the ciliary processes between the choroid and the iris) exerts a tension upon the lens which keeps it flattened, particularly along the anterior surface, but when the eye becomes fixed on near objects the ciliary muscle contracts and in doing so draws forward the choroid coat and thus relaxes the tension of the ligament, in consequence of which the anterior surface of the lens becomes more convex.

power of adapting themselves in this way to view near objects or when the eyeball is naturally shallow, light rays would have to extend behind the retina to come to a focus and the vision of objects near at hand, especially fine ones as the letters of printed matter, is blurred and indistinct. This condition is known as *hypermetropia* or *far-sightedness.* The term *far-sighted* is applied to eyes with this defect because when it is not very great the individual can often see things at a distance better than people whose eyes are normal. Defective vision due to this condition is corrected by wearing glasses, by means of which the rays are sufficiently refracted to make up for the defective structure of the eyeball or action of its refractory apparatus, as the case may be. Failure of accommodation, due to old age, is known as *presbyopia* or *old-age sight.*

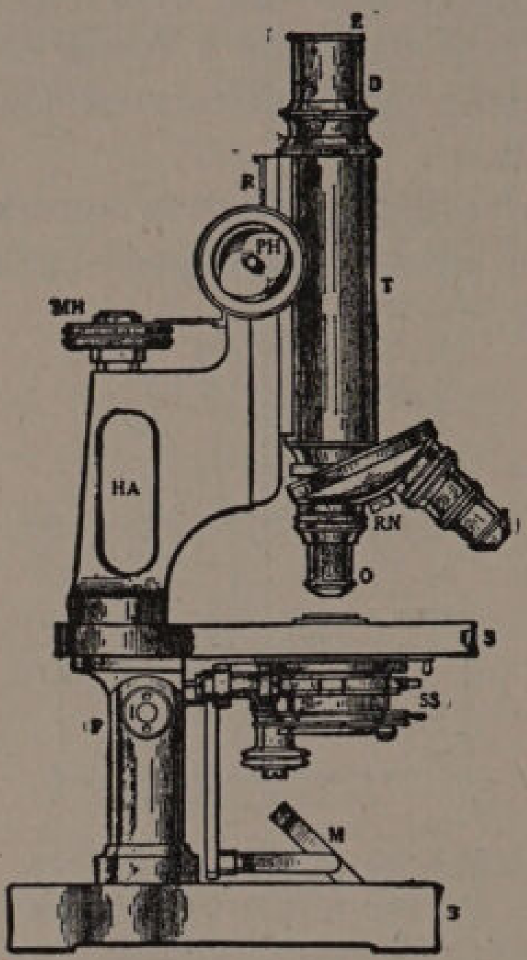

FIG. 47. MICROSCOPE.
e, Eyepiece. d, Draw tube. t, Body tube. rn, Revolving nosepiece. o, Objective. ph, Pinion head. mh, Micrometer head. ha, Handle arm. s, Stage. ss, Substage. m, Mirror. b, Base. r, Rack. p, Pillar. i, Inclination elevation of handle arm.

Another defect of the eyes that interferes with proper refraction is that known as *myopia* or *near-*

sightedness. This is usually due to a too great curvature or too great depth of the eyeball, in consequence of which light rays come to a focus before they reach the retina. To obviate the blurring of vision due to this condition concave lenses—*i. e.*, those thinner in the center than at the edges—are worn, because they cause rays of light to diverge before entering the eye and thus it takes them longer to come to a focus afterward.

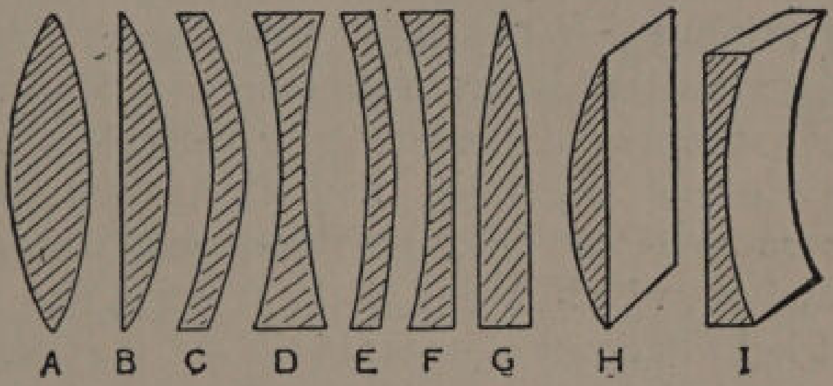

FIG. 48. DIFFERENT TYPES OF LENSES.

A, Double convex. B, Plane convex. C, Concave convex. D, Double concave. E, Convex concave. F, Plane concave. G, Prism convex. H, Cylindrical convex. I, Cylindrical concave.

Light is diverged by a concave lens because the middle of the wave is retarded less than the edges, as the lens is thinner in the center than at the edges.

Astigmatism.—Still another cause for defective refraction in the eyes is the condition known as *astigmatism* in which the curvature of the refractory apparatus is not equal in all parts. This makes it necessary for the ciliary muscles to be constantly in action while the individual is reading, focusing for the lines of letters that run in different directions. This soon results in eyestrain. Cylindrical lenses, which have a plane surface in one axis and a curved surface in the axis at right angles to it, are worn to correct the condition.

Prism glass.—Glass with a rough surface will be often seen in windows that open on to a court or narrow street or passage. This is called *prism glass*. It improves the light in a room because the prism-like projections on its surface refract the light waves passing through it, causing them to penetrate farther into the room than they naturally would.

The rainbow.—When light waves enter drops of water or glass that is cut in the manner of the beveled edge of a mirror or of a prism, these of different lengths suffer unequal refraction and thus they become separated and the different colors are discernible. The rainbow, which is due to the fraction and reflection of light in drops of water in the air, is an example of this kind of refraction of light waves.

Polarization

Nature of polarization.—By polarization is meant making light vibrations take place in one direction. The streak of regularly reflected light striking back from a mirror is an example.

How light can be polarized.—Light is polarized by reflection from a flat, highly polished surface, as a mirror, by refraction and by passage through such substances as the mineral tourmaline and iceland spar. As these substances cause polarization, it is thought that their electrons must vibrate in one particular direction and that only waves that will vibrate in the same way can pass through them.

Light is often polarized from a mirror in order to thoroughly illuminate body cavities. It is polarized by a special apparatus known as a *polariscope*, in pathological work, for the detection of glucose in fluids, such as the urine, for when polarized light falls upon glucose it is rotated to the right and the degree of rotation will depend upon the amount of glucose present and thus it is an easy and accurate way of determining the quantity of glucose in a solution. Levulose, on the contrary, turns polarized light to the left. Also, there are a number of other substances that turn it to either the right or left in varying degrees and thus it is used in many varieties of scientific work, especially as a means of identifying substances which rotate it.

Phosphorescence and Fluorescence

Phosphorescence.—Certain substances, after being exposed to a strong light, will emit light in the dark. This property is termed *phosphorescence* and matter which possesses this power is said to be *phosphorescent*. Paints which are luminous in the dark contain some phosphorescent material, usually barium chlorid. A certain class of bacteria, called *photogenic*, produce this quality in matter, especially fish, in which they produce decay.

Fluorescence.—Certain substances, such as the mineral fluorite, fluorescin, platino-cyanid of barium, uranium glass, and solutions of quinin sulphate, have the power to change the length of ether waves, and thus of changing heat waves to light waves and *vice versa* to a marked degree, and of reflecting color waves different from those thrown upon them. Substances which do this are said to be *fluorescent*. This name was given, because the mineral fluorite was the first substance in which the property of fluorescence was noted.

CHAPTER VII

ELECTRICITY

Theories Regarding the Nature of Electricity and Electrification —Conductors and Non-Conductors—Different Methods of Generating Electric Currents—Nature and Action of Chemical Cells and Batteries for Generating Currents—Electrolysis—Electroplating.

Derivation of name.—The term electricity was derived from the Greek work *electron* = amber, because the first record of electrization was in connection with this substance. This record, which was made by the Greeks, some 500 years before Christ, was to the effect that if amber was rubbed with a dry cloth it would attract light objects to itself.

Experiment 11. Object: To study electrification.

Procedure: (a) Rub a glass rod for a few minutes with a dry woolen cloth or a piece of silk and then drag it through some small pieces of paper scattered on the table.

If the rod is rubbed long enough, under suitable conditions, see page 107, it will attract the pieces of paper towards itself.

(b) Hang this rod by a piece of silk thread to a wire bar or stirrup, electrify another glass rod in the same manner as the first rod and hold it against the free end of the suspended rod, the latter will swing away from it; *i.e.*, it will be repelled.

(c) Rub a piece of sealing wax with silk and hold it near the suspended glass rod, the latter will swing toward it; *i. e.*, it will be attracted.

(N. B. Be careful while doing this experiment not to let the part of the rod that is to be used come in contact with the hand or other object. The reason for this will be seen later in the section on conductors.) In the same section it will be also explained why this experiment is seldom successful in wet weather.

There are two things to be learned from this experiment: (1) Friction will produce electrification. (2) Substances of similar material electrified in the same way repel each other, but electrified objects of different material or of the same material differently electrified may attract each other. This is because there are two kinds of electrification—*i. e.*, of electric charges—and one of the laws of electrical action is: *Like charges repel each other and unlike charges attract.* The two kinds of charges have been arbitrarily named *positive* and *negative*, and a positively electrified body is often defined as: *An electrified body which acts toward other electrified bodies as does a glass rod that has been rubbed with silk.* Likewise a negatively electrified body is said to be: *One which acts toward other electrified bodies as does a stick of sealing wax which has been rubbed with flannel.*

But one kind of electricity.—Though the expressions *negative electricity*, *positive electricity*, *static electricity*, *current electricity*, etc., are constantly used, there is but one kind of electricity. For this reason, it has been suggested that the terms *negative* and *positive electrification* or *negative* and *positive* *electric charges* be used instead of negative and positive electricity. By static electricity is meant, that which is not flowing, as, for example, when it is passing over a wire. In the latter case it is known as *current electricity*, or, more correctly, an *electric current*. The difference, however, lies, not in the electricity, but in the manner by which it is made manifest.

Some theories regarding the nature of electricity.—Though a sufficient degree of knowledge, of the ways in which electric charges and currents can be produced and of the results of their action, has been obtained to allow of electricity being applied in a great many ways to the service of man, its nature is still a matter of conjecture. Some theories that have been advanced are as follows:

The two-fluid theory.—When the ease with which electricity moved over certain substances was first realized, the idea was conceived that it was a fluid. Or, rather, it was thought at first that there were two fluids, one giving rise to positive and the other to negative charges.

The one-fluid theory.—About the year 1750, Benjamin Franklin advanced what is called *the one-fluid theory*, which holds that there is but one kind of fluid and that positive charges are the result of an excess of electricity and negative charges of a lesser amount. Because of this supposition, Franklin used the plus sign (+) as a symbol for positive charges, and the minus sign (—) as a symbol for negative charges. These signs are still used, though, according to the electron theory, which many scientists think is probably nearer the truth, a positively charged body has lost, instead of gained, electrons.

The electron theory.—The electron theory is based on the assumption that the atoms of all elements are composed of electrons. It holds that each atom of an element consists of inconceivably small *positive corpuscles* that are surrounded by thousands of still smaller bodies called *negative electrons*. These negative electrons are, it is thought, constantly revolving with exceedingly great rapidity around the positive corpuscles and were it not for the attraction of the positive corpuscles, they would, it is thought, fall away.

Ordinarily, the majority of substances are in a state of electrical equilibrium—*i.e.*, neither positive nor negative—but, it is assumed, some of the more easily detached revolving electrons may be transferred from one atom to another without dissociating the atom; the only change being that the atoms of the substance by losing some of their negative electrons have their positive charge strengthened and the substance is then said to have an *electro-positive charge*. On the other hand, the substance that gained the negative electrons has its negative charge strengthened and is said to be *electro-negative*. This is what it is thought happens in Experiment 11. The friction of the silk on the glass caused the atoms of the latter to part with some of their negative electrons to the silk and thus the glass became positively and the silk negatively charged. On the contrary, the silk gave some of its electrons to the sealing wax and thus it became positively and the wax negatively charged.

The reason why the silk and other material gain electrons in one case and lose them in another is thought to be due to some difference in the number and arrangement of the electrons by which certain substances are able to take electrons more readily from some materials than from others.

It is believed that positive corpuscles are never dissociated from the atoms.

Nature of an electric current.—As stated on page 104; an electric current is simply a flow of electricity through some object and this flow, it is now believed by many scientists, consists simply of a passing of negative electrons from atom to atom, usually in one direction, along the object.

Conductors.—There is a great difference in the ease and rapidity with which an electric current passes through different kinds of matter. For example, copper wire will conduct electricity a million times better than water. Substances through which a current passes readily are said to be *good conductors*. Those through which it does not pass readily are called *poor conductors*, and very poor conductors are termed *non-conductors* or *insulators*. No substance however is a perfect conductor nor a perfect insulator.

The following list shows the relative conducting value of some common substances. They are mentioned in the order of their conducting value beginning with the most efficient. The last nine are considered insulators.

Silver
Copper
Aluminium
Zinc
Iron
Carbon
Dilute sulphuric acid
Water
Human body
Linen
Cotton
Wood
Dry air
Silk
Sealing wax
Glass
Hard rubber
Porcelain
Shellac
Paraffin
Oils

Moist air is a much better conductor of electricity than dry air, which is why Experiment 11 is not likely to be successful when the humidity is very high. Dry air, being more rarefied than moist air, is a very poor conductor, thus in dry air a substance retains its charge for some time. A common example of the favorable influence of dry air on electrification is seen also when brushing the hair in cold, dry weather. For the crackling sound that occurs and the flying of the hairs apart are due to electrification; and that the hairs fly apart, but cling to the brush, shows that like charges repel each other and unlike charges attract each other. Glass, sealing wax, silk, etc., were used for the experiments because, being poor conductors, they retain their charge for some time. Metals would not serve such a purpose, since, being good conductors, they lose their electrification unless it is being continuously produced as for an electric current. If the hand is rubbed over the electrified rod, the latter will lose its charge, which will be imparted to the hand, and the same thing will happen if the rod comes in contact with metal or any substance that will act as conductor.

How an electric current is produced.—In order to start and maintain an electric current along a conductor some external force is necessary. This is generally obtained by the use of an electric cell or battery or of a dynamo. The force obtained by use of a cell or battery is the result of chemical action, and that obtained with a dynamo is, as will be seen later, due to some form of mechanical action or of heat.

Chemical batteries.—Chemical electric batteries consist of two or more cells and the essential parts of a cell are:

A jar
Two electrodes
An electrolyte
A conductor

The electrodes.—These are two flat plates or bars of metals or other substance by which the current enters and leaves the electrolyte. The electrode at which the current is started is known as the *anode* (from the Greek *an* = in and *hodos* = a road) and the electrode by which the current has been supposed to leave the cell is known as the *cathode* (from the Greek *cath* = away and *hodos* = a road).

Zinc is very generally used for the anode of battery cells for two reasons: (1) It is easily acted upon by liquids that are convenient to use for electrolytes, and this chemical action provides the energy necessary to procure the electric current; (2) As zinc is dissolved by the acid its atoms part with their negative electrons, leaving them with the undissolved metal. The necessity for this will be seen later. Since the anode thus becomes negatively charged, it is spoken of as the *negative electrode* and is represented by the minus sign (−).

A cathode for a chemical cell is generally made of either copper or carbon; these substances not being acted upon by electrolytes that dissolve zinc. As will be seen later, when a cell is in action, the cathode acquires a positive charge and it is therefore often referred to as the *positive electrode* and it is represented by the plus sign (+).

Electrolytes and electrolysis.—An electrolyte is a solution that conducts and is decomposed by electricity. The process that occurs in an electrolyte is spoken of as *electrolysis*.

The electrolytes generally used for chemical batteries are sal ammoniac and sulphuric acid, but solutions of any inorganic acids, bases, and salts will undergo electrolysis.

Nature of electrolysis.—It is thought that when inorganic acids, bases, and salts are dissolved in water a portion of their molecules become separated into two or more parts, called *ions*, which move about in the solution independently of each other and of other atoms and molecules (hence their name, which is derived from a Greek word meaning *wanderer*). Ions, are unlike other atoms and molecules in that they have an electrical charge; in fact, a common definition for an ion is *an atom or radical having an electric charge.*

When a substance undergoes electrolysis, one of the ions resulting from the dissociation will have a positive and one a negative charge. According to the electron theory, this is because the atoms of the element or elements forming one ion gave up some of their negative electrons to the other ion in the process of decomposition. The ion losing the electrons is, of course, the positive ion and the one gaining electrons, the negative ion. When sulphuric acid (H_2SO_4) is thus decomposed, the hydrogen ions are the positive ions and are said to have or carry a positive charge and the sulphate ion (SO_4) has a negative charge. When sodium chlorid ($NaCl$) is dissociated, the sodium (Na) has a positive charge and the chlorin (Cl) a negative charge; when sodium hydroxid ($NaOH$) is dissociated, the sodium (Na) has a positive and the hydroxyl *ion* (OH) a negative charge, and so on.

Positive ions are called *cations*, because they go to the cathode and, like a positive charge, they are represented by the (+) plus sign. Negative ions are called *anions*, because they collect at the anode and they are represented by the (−) minus sign.

Arrangement of a simple cell.—The simple cell, often used in laboratories for the purpose of observing the action that occurs in chemical batteries, is known as a *Galvanic* or *Voltaic cell;* because Galvani, an Italian anatomist, in 1786, was the first to discover that a continuous current of electricity could be procured by chemical means, and Volta, an Italian physicist, in 1800 invented an arrangement similar to this form of cell. It is arranged as follows: A strip of zinc and a strip of copper, each having a piece of copper wire soldered, or otherwise attached to their upper end, are fixed, one on either side, in a jar containing dilute sulphuric acid solution. The electrode, electrolyte, and conductors are known as the *electric circuit*, since it is through these things that the electric current passes. When the wires attached to the electrodes are joined, the circuit is said to be *made* or *closed* and when the wires are disconnected the circuit is said to be *open* or *broken*.

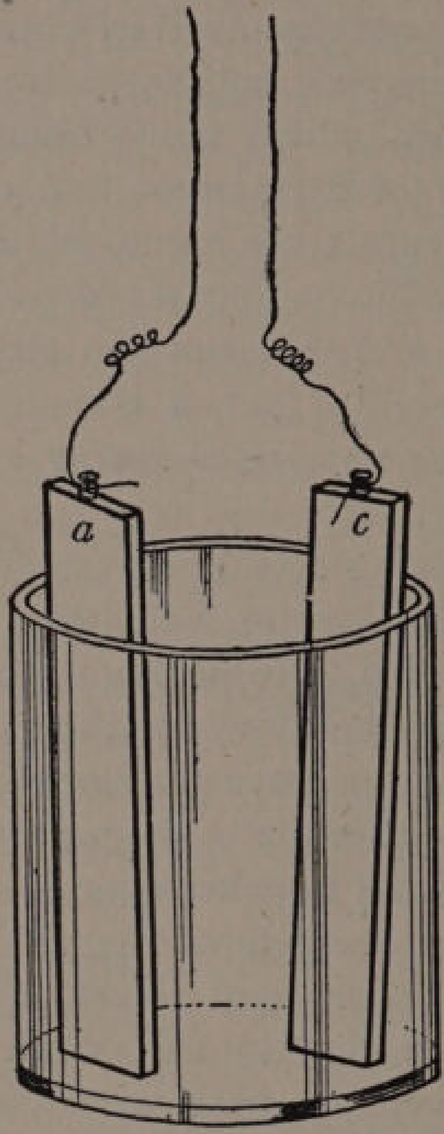

FIG. 49. VOLTAIC CELL.
(*a*) Anode or positive pole −.
(*c*) Cathode or negative pole +.

Chemical action that occurs in a cell.—When the zinc (Zn) and the sulphuric acid (H_2SO_4) come in contact, the zinc begins to dissolve gradually and the H_2SO_4 to split into H and SO_4 ions. The atoms of the dissolved zinc thereupon unite with the sulphate ions, forming the salts known as *zinc sulphate*, and the hydrogen appears in the form of bubbles first on the zinc strip and, after the circuit is closed by connecting the wires, on the copper plate or cathode.

Theories of electric action.—The manner in which the electric current passes through the circuit has not been as satisfactorily ascertained as the nature of the chemical action and, at present, there are two theories regarding this. They are about as follows:

Theory 1. As the zinc goes into solution in the acid, the atoms separated from it leave some of their negative electrons behind in the undissolved portion of the metal; thus, this soon acquires a negative charge, and the solution surrounding it, which contains the dissolved zinc, becomes positively charged. This part of the solution repels the hydrogen atoms, which are also positively charged, and, consequently, they fly to the copper pole. Of course, as already stated, the zinc unites with the SO_4 and becomes electrically neutral—*i. e.*, neither positive nor negative, but, while the circuit is closed, there is always enough ununited positive zinc in the solution to repel the hydrogen atoms. When the hydrogen ions reach the copper they part with their charge and collect as bubbles of hydrogen gas upon it. As the copper electrode receives the charge from the hydrogen ions it becomes positively charged. The electric current thus started passes up the copper plate over the wire conductors to the zinc electrode, down this to the electrolyte, and back to the copper electrode, and so on.

Theory 2, which has been formulated since the electron theory of the nature of electricity was advanced, considers that the current leaves the cell from the zinc pole or anode and not, as has always been supposed, from the cathode.[1] It holds that, due to the force generated by the chemical action, the negative electrons left behind on the zinc electrode by the atoms of the metal which goes into solution in the acid, as described in Theory 1, fly up the electrode and over the wires to the cathode, and so on.

If the two wire conductors are disconnected, thus breaking the circuit, chemical action in the cell soon ceases. The reason for this, it has been assumed, is that, when the current ceases, the zinc acquires such a strong negative charge that it holds on to the positively charged zinc atoms and keeps them from going into solution, and, at the same time, the copper pole acquires such a strong positive charge that it ceases to take the positive charge from the hydrogen. As soon, however, as the circuit is closed, the flow of the current is reëstablished and the chemical action proceeds as before.

The wire conductors connecting the two electrodes need not be short, as in the cells used for demonstrations in the laboratory but for example the cells of a battery are often put in the basement of a house and the wires extend from this to a plate behind the bell push button at the front door. One of the wires, on its way to the front door, extends through the kitchen or other room where it is connected to the bell apparatus. The essential parts of the latter

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[1] *Scientific Ideas of To-day*, Charles R. Gibson, Seeley, Service & Co., Chapter v.

are: a bar of soft iron which is attached to the hammer that strikes the bell, and another iron bar, called an *electromagnet*, that is in connection with the wire conductor and becomes magnetized when the electric current passes through it.

Each of the two wires at its terminal behind the push button is attached to a small strip of metal. This is where the circuit is open and closed. Ordinarily, the metal strips do not come in contact and the circuit is open, but when the button is pressed the metals come in contact and the circuit is closed and the current at once passes through the wire. As it passes through the *electromagnet* it, as previously stated, magnetizes it, whereupon it, since magnets attract iron, draws forward the bar to which the bell-hammer is attached and causes it to strike the bell. So long as the button is pressed and the circuit thus kept closed, the current passes through the circuit and this type of a bell continues to ring. As soon, however, as pressure on the button is released, the circuit is opened and the bell ceases to ring.

Electric batteries.—The current generated in one cell would not be sufficient to force its way over a long conductor, such as described in the preceding paragraph, and therefore, for such purposes, a battery, consisting of two or more cells is used. The different cells are connected by a wire extending, usually, from the cathode of one cell to the anode of the other.

The principle of the cells of batteries in common use is the same as the simple cell just described, but carbon is often used for the cathode and sal ammoniac for the electrolyte. Sal ammoniac is usually preferred to sulphuric acid because the latter, unless some means is taken to prevent it, causes polarization; *i.e.*, the hydrogen gas which collects on the cathode, being a poor conductor of electricity, prevents the passage of the current, according to Theory 1, to the cathode and, according to Theory 2, from the cathode. If the current is stopped in this way, the results will be the same as though the circuit were broken, and the battery is useless until the condition is rectified.

Dry cells.—The so-called *dry cells* (they are moist rather than dry) are used for so many purposes that require only a small current of electricity that a few words of description may not be out of place. A dry cell consists of a zinc cylinder (this acts both as a case for the cell and as the anode) lined with a moist paste made of strong sal ammoniac and plaster of Paris or other porous substance and containing in the center a rod of carbon and manganese dioxid. The top of the cell is sealed with wax or pitch to prevent evaporation of the moisture from the paste. There is a wire connected to the zinc cylinder and also one to the carbon rod. Two great advantages of dry cells for many purposes are that they are comparatively small and contain no liquid to be spilled. The action of a pocket flash lamp is obtained by means of one or more small dry cells, which are contained in its cylinder.

Storage batteries.—Storage batteries or, as they are also called, secondary batteries or accumulators are much used for operating telephones, fire-alarm circuits, furnishing power for electric automobiles and the like. Also they are used in power houses to store excess electricity generated by the dynamo, thus providing a reservoir to draw upon when an extra amount of current is required.

A common form of storage cell consists of perforated plates of lead, the perforations of which are filled with a paste of red lead and litharge mixed with sulphuric acid. These plates are kept in a suitable receptacle, covered with sulphuric acid. They are charged with electricity by sending a current through them from an outside generator—*e. g.*, a dynamo. This current dissociates the H_2SO_4 and the hydrogen ions move to one set of plates and change the paste in them to a spongy metallic lead. These plates then act as a cathode. The SO_4 ions, at the same time pass to the opposite plates and change the paste in them to lead peroxid. This set of plates acts as an anode.

When the circuit is closed, some of the lead peroxid is decomposed and gives rise to an electric current, just as occurred in the galvanic cell.

These batteries are known as *secondary batteries*, because a primary current is used in their manufacture. This name is really more appropriate than that of *storage batteries;* for these batteries do not store electricity, but rather the electric energy has been changed to latent energy, which, when the circuit is closed, is, by chemical action, transformed once more to electrical energy. This action is analogous to that which occurs in plants and coal, as described in Chapter II. When the peroxid of lead has been all used, the battery can be recharged by allowing a current of electricity to pass through it as when it was made. Thus these batteries can be used for a long time.

Electroplating.—Some of the cheaper metals are often coated with more expensive metals by means of a process known as *electroplating*. This process, though different in some respects, is similar to that which occurs in a voltaic cell. The electroplating of spoons with silver is a good illustration of the method. The apparatus used for this purpose consists of a tank across which copper rods are placed. These rods, which act as conductors, are connected with wires attached to the battery. The spoons are hung from the rod connected with the wire from the cathode of the battery, and bars of silver are suspended from the bar connected to the anode wire. Enough electrolyte is used to cover the spoons and silver bars. The electrolyte in silver plating consists of silver salts, such as silver nitrate. When the electric current is turned on, the silver salts are decomposed, the freed silver passes to the spoons, which thus become coated with the metal. At the same time, the silver bars, which act as anodes, are being slowly dissolved in the solution and, uniting temporarily with the nitrate ions, they provide fresh salts.

If a gold plating is wanted, gold bars will be used for the anode and gold salts for the electrolyte. For a copper coating, copper bars will be used for the anode and copper salts for the electrolyte.

The chief difference between the action in electroplating and that in a galvanic battery is that, in the former, the electric current is essential for the dissolving of the anode, since there is no chemical action between it and the electrolyte, and in the galvanic cell, the chemical action between the electrolyte and the anode starts the electric current.

CHAPTER VIII

ELECTRICITY AND MAGNETISM

The Dynamo—Different Kinds of Magnets—Magnetism—Induction Coils—Transformers and Other Electrical Appliances—Electric Currents as a Source of Heat and Light—Measurement of Electricity—Static Electricity—Physiological Action of Electricity—Cathode Rays—X-Rays—Radio Rays

The Dynamo

THE electric current produced in a chemical cell or battery is not suitable for operating lighting and heating appliances, motors and the like. For one reason, the current would not be sufficiently strong, and, for another, the anode and electrolyte of a battery would have to be constantly renewed, if the circuit were kept closed for hours at a time. It has, therefore, been only since the invention of the dynamo that it has been possible to use electricity to any extent for such purposes.

The dynamo is an apparatus in which the mechanical energy from a steam- or water-driven engine, or the heat energy developed in engines fed with gas, gasoline, or similar fuel, is transformed into electrical energy. In other words, as the result of the force of steam or water power or of heat, the dynamo is started and kept in action and this action gives rise to an electric current which, according to the electron theory, is a streaming of negative electrons along the wire conductor leading from and to it.

Magnets

Magnets.—One of the essential features of dynamos, electrical motors, and many other electrical appliances being an electromagnet, it seems well to say a word here about magnets, magnetism, and electromagnets.

Natural magnets.—In various parts of the earth, there is found an iron ore that will attract small particles of iron and steel.[1] Such ore has been named *magnetite* (because it was found first near Magnesia in Asia Minor), natural magnet, and lodestone, meaning leading stone.

Artificial magnets.—When iron and steel are rubbed on a natural magnet, or on steel that has been rubbed by a natural magnet, they will likewise attract steel and iron. Metals that have been thus treated are called *artificial magnets*. Cobalt and nickel also can be magnetized to a slight extent, but iron and steel are the only substances that assume strong magnetic properties.

Magnetic properties.—The chief properties or characteristics of magnets are: (1) Induction, *i. e.*, the power to induce magnetic properties in unmagnetized steel and iron, even at a distance. (2) Magnetic attraction, *i. e.*, the power to attract iron and steel. (3) Polarity.

Polarity.—If a long narrow magnet, either natural or artificial, is suspended by a fixture at its center in a manner that will allow it to turn freely, it will be found

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[1] As will be seen in Chapter XVI., steel is a form of iron.

that the magnet will always come to rest pointing approximately north and south. Also, it will be found that when a magnetized bar is put in the midst of iron filings, the latter will collect around the two ends of the bar and few if any filings will attach themselves to the sides. If, however, the bar is broken in the center, the broken ends will attract the filings. The ends of a magnet have been named its *poles* and the two poles are known respectively as the *north-seeking* or *north pole* and the *south-seeking* or *south pole.* This property of forming poles is spoken of as *polarity.*

Just as electrically charged bodies repel those with a similar charge and attract those with an unlike charge, so the north pole of one magnet repels the north pole of another magnet and attracts its south pole, and *vice versa.*

Why magnetite was called lodestone.—The earth is a magnet and, like all magnets, it has its north and south poles. The north-seeking pole of the magnet was so named because it is attracted by the north magnetic pole of the earth, and the south-seeking pole, because it is attracted by the south magnetic pole. Thus, the north pole of the long narrow magnet, constituting the needle of a mariner's compass, points toward the north magnetic pole, and the degree of inclination of the needle shows the direction in which the ship is going. It was for this reason that magnetite was named *lodestone*, which means *leading stone.*

Cause of magnetism.—It is thought that each molecule of iron and steel is a magnet at all times, but that, ordinarily, the molecules are not lying in any definite position and that, therefore, one molecule neutralizes the effect of the others. When the iron is stroked by a magnet, however, all these small molecular magnets are forced to turn with their north poles all in one definite direction, and when the molecules are thus placed their combined effect is strong enough to act.

This theory accounts for at least three truths about magnetism: (1) Only the poles of a magnet show magnetic attraction (naturally this would be the case when the molecules are in position, for then the south poles of one line of molecular magnets are in apposition with the north poles of another line, and thus they neutralize each other); (2) if the bar is severed, the severed ends act as magnets; (3) if a magnet is heated to red heat, it will lose its magnetic properties; presumably, because the heat increases molecular movement to such an extent that the molecules lose the position they were made to assume by the magnet, and they then once more neutralize each other.

Experiment 12. Object: To study magnetic fields and lines of force.

Procedure: Sprinkle some iron filings upon a piece of paper and place this over a magnetized bar. Then gently tap the paper. The filings will arrange themselves in curved lines extending from one pole to the other, these lines will run side by side as seen in Fig. 50, and will not cross each other. This is because the magnetic influence is not confined to the immediate vicinity of the magnet, and the filings coming within the field in which the magnetic force extends become magnetized. The area in which the magnetic force is felt is spoken of as the *magnetic field* or the *field of force*, and the attraction which tends to arrange the iron filings in the curved lines is spoken of as the *lines of force*.

This property of magnets is of great value, for the action of the electromagnets in dynamos, motors, induction coils, and other kinds of electrical apparatus is largely dependent upon it.

It was Michael Faraday, an English physicist, who, about the middle of the nineteenth century, discovered that a current of electricity could be induced in a wire by moving it in a magnetic field, and who devised the appliance that was the forerunner of the dynamo and many other electrical devices in use at the present time.

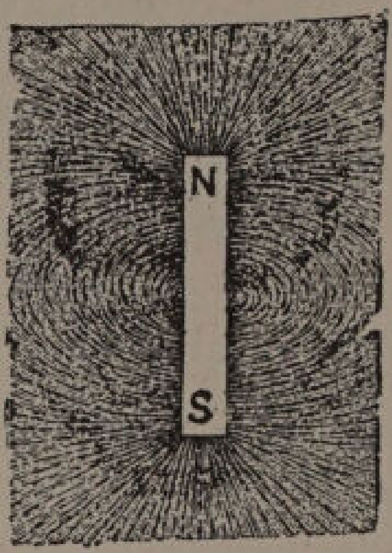

FIG. 51. IRON FILINGS ON PAPER OVER BAR MAGNET, SHOWING MAGNETIC FIELD AND LINES OF FORCE.

Electromagnets.—Soft iron is more easily magnetized than steel, but the latter maintains its magnetic properties for a long time, while the former soon loses them. Soft iron is very readily magnetized by electricity, but it ceases to be a magnet as soon as the current stops. Iron magnetized in this way is called an *electromagnet*. A common type of electromagnet consists of a core of soft iron or iron wire wound with insulated wire. The iron becomes magnetized by the current passing through the wire.

A simple form of electromagnet and the method of its action were referred to in the description of the electric bell, page 113. Another similar form of electromagnet is that inserted in the framework of the front door in some apartment houses, in such a way that, when it is magnetized by the closing of the circuit, it draws the door catch back, allowing the door to be opened. The circuit is closed by the pressing of a button in an apartment. A large number of electrical appliances and motors are operated in much the same way.

Galvanic and faradic currents.—By a galvanic current or galvanism is meant a current obtained from a chemical cell or battery; by a faradic or induced current is understood a current that is induced in a conductor by the action of a magnet or an electrified body that is not connected to it.

The current produced by dynamos and induction coils are examples of induced currents. Induced currents are known also as *secondary currents*. Induced currents that flow first in one direction and then in the other along the secondary coil are called *alternating currents*. Those that flow in one direction only are called *direct currents*.

Induction coils.—The essential parts of an induction coil, see Fig. 52, are a core of soft iron (C) around which are wound a few turns of coarse copper wire. This wire, which is known as the *primary coil*, is connected into the circuit of a battery (B). At its connection with the battery, it is in contact with a fine steel spring (D), at the end of which is a piece of soft iron. This acts as a current interrupter. Over the primary coil is wound what is known as the *secondary coil*. This consists of a great many turns of fine copper wire. The ends of the secondary coil terminate in what are known as the spark points (SP). When the circuit is closed, the current flows through the primary coil, magnetizing the iron core. This attracts the spring, pulling it back, which breaks the circuit. The core thereupon loses its magnetism, the spring falls back, and the current is remade. This cycle is repeated so long as the coil is in action, thus the current is being continuously interrupted, though only momentarily. This is of importance in the action of the apparatus. The magnetic field created by the current in the primary coil induces a current in the secondary coil which, as the wire composing it is much longer and finer than that of the primary coil, is much stronger than that in the latter. The reason for this will be seen later. When the current is sufficiently strong, spark points pass be-

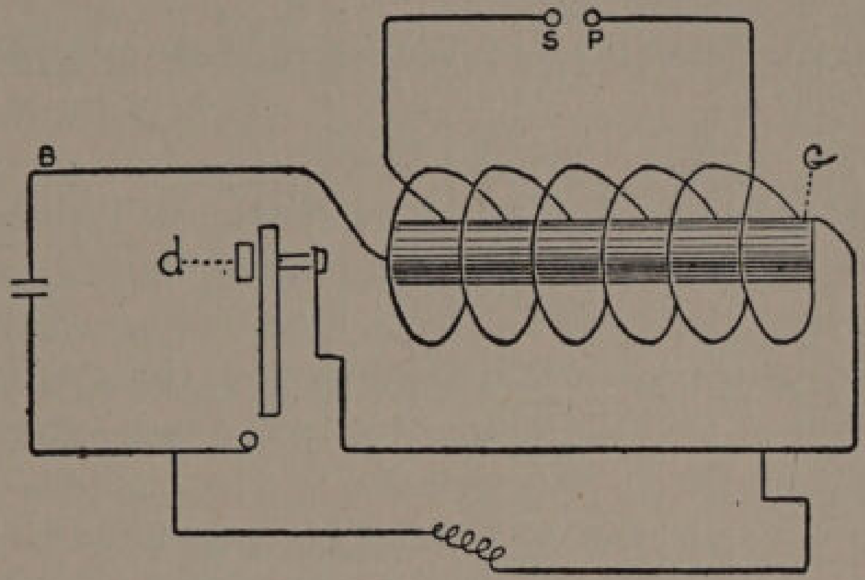

FIG. 52. DIAGRAM SHOWING PART OF AN INDUCTION COIL.

(B), Battery.
(C), Primary coil.
(D), Spring.
(S, P), Spark points.

tween the secondary terminals (S, P). Induction coils form essential parts of many kinds of electrical apparatus. The various varieties of faradic induction machines used in electrotherapy are modifications of it.

Electric Currents as a Source of Heat and Light

Heat effects of electric currents.—Whenever an electric current passes along a wire conductor, the latter becomes heated. The degree to which it will do so depends upon two things—viz., the strength of the current and the amount of resistance offered by the conductor to the passage of the current. The amount of resistance will depend largely upon: (1) The kind of wire used, some metals being poorer conductors and thus offering more resistance than others. (2) The length of the wire; the longer the wire, the greater the resistance. (3) The circumference of the wire; the greater the circumference of the wire, the easier it will be for the current to pass and thus the less will be the resistance. (4) The temperature of the wire; less resistance is offered by a cold, than by a hot, wire.

Since resistance to a current causes heat, wires that will offer the least resistance are used for service wires and those intended to carry a current wanted to produce mechanical or chemical results; since, if heat is not required, to have any unnecessary energy expended as heat means a loss, as there will be then just that much less energy left to accomplish the work for which the current is intended. On the other hand, where heat is required, means are taken to secure a varying amount of resistance, the amount depending upon the intensity of heat required. For instance, in the electric furnaces used to reduce ores, material is used that will offer tremendous resistance to the current, and an excessively high temperature, 3000° C. and over, is attained. Of course, the greater the resistance provided, the stronger must be the current used. Iron and platinum wire will offer more resistance than copper and will therefore become hotter.

The heating element of an electric flatiron usually consists of a coil of fine wire. This is placed within a hollow iron casing which it heats. Electric warming pads consist of coiled wire enclosed in asbestos cloth and covered with flannel. Electric stoves are heated by a coil of wire embedded in enamel that is baked to the inner side of the metal serving for the top. In electric broilers and toasters the heat is radiated from coiled wire fastened to the framework. The various other heating appliances in common use are made in these or similar manners, the heating principle of the majority being some kind of fine wire that is so coiled that a long piece can be used in a small space.

Electric lights.—The two common forms of electric lights are the incandescent electric lights and the arc light.

Incandescent lights.—Since a high degree of heat is necessary to produce light, two requirements of the wire filaments used as the lighting element of incandescent lights are that they offer a high degree of resistance to the current and that they are melted only at temperatures considerably above that at which they become incandescent. Tungsten, tantalum, and a carbon filament—made by carbonizing a special kind of cotton thread—answer these requirements. Until lately carbon filaments were almost universally used, but the metals tungsten and tantalum, though more fragile and expensive, stand a higher temperature than the latter and can be arranged so that they will give a stronger light without the use of any extra current. They are, in fact, more than twice as efficient as the carbon lamp. Filaments of both these metals offer less resistance than carbon and therefore it is necessary for them to be longer than those made of carbon, and they are therefore differently arranged as can be seen by comparing a tungsten or mazda lamp with one that has a carbon filament. The longer the filament, the brighter the light will be. The lighting filaments are always encased in a glass globe from which all air has been exhausted, for if they were heated in the air they would become oxidized and thus rendered useless for their purpose. The two ends of a filament are connected to platinum wires which are sealed into the small glass tubes that project from the screw into the bulb. Platinum is used for this purpose because it is one of the few substances that is contracted and expanded at the same rate as glass by changes in temperature, and the glass would be broken were it in contact with a substance that did otherwise. The platinum wires are separated from each other by insulating material and are connected in the screw to separate pieces of metal. The screw fits into a socket provided with similar pieces of metal to which the two service wires (that from the anode and the cathode of the electric machine) are connected. As the metals of the screw and lamp come in contact when the switch is turned, the circuit is then closed and the current flows through the lighting filament, which is almost instantaneously heated to incandescence.

Arc lights.—The lighting principle of the arc light which is used for the lighting of streets, large buildings, and other purposes requiring a powerful light, consists of two carbon rods. These are in contact when the current is turned off, but as soon as it is turned on it passes through an electromagnet which magnetizes an iron bar from which the upper carbon is suspended and this pulls the carbons about three-eighths of an inch apart leaving a space that is known as the *arc*. At the same time, the upper carbon is heated to a white heat and, when this happens, it gives off incandescent particles which jump across the arc. The lower carbon also becomes heated, but the upper carbon always serves as the anode, the current passing from it to the lower one and out. The incandescent particles, likewise, pass from the anode to the cathode and, consequently, the upper carbon becomes dented in the center and the lower one becomes cone-shaped. Both carbons wear away, but the upper one more quickly than the lower. As this happens the gap between the two increases in size and they must be brought nearer together. On street lamps this is usually done automatically by the action of electromagnets. In projectoscopes and similar kinds of apparatus the regulation is usually done by hand.

Meaning of Terms Used in Connection with the Measurement of Electricity

Electromotive force.—This term is used to signify the pressure of electricity. A powerful current will have a high electromotive force and a weak one a low electromotive force.

Unit.—In connection with measurement, this word is used to denote a quantity assumed as a standard of measurement.

Volt.—The volt (so called after Alessandro Volta, an Italian physiologist and physicist, 1745–1827) is the unit of electromotive force, it is the force necessary to cause one ampere of current to flow against 1 ohm of resistance.

Ordinary galvanic and storage cells furnish between 1 and 2 volts; batteries may furnish 20 or more volts, according to the number of cells. A very much higher voltage can be obtained by use of the dynamo. That employed on the trolley-car circuit is about 550 volts; that furnished houses is between 110 and 220 volts.

Ohm.—The ohm (named after George Ohm, a German electrician, 1781–1854) is the unit of electric resistance. As stated on page 106, even the best conductors offer some resistance to an electric current.

Ampere.—The ampere (called after André Maria Ampère, a French physicist and mathematician, 1775–1836) is the unit of electro-current strength. It is the amount of current afforded by one volt of electromotive force against one ohm of resistance. In other words amperes denote the quantity of current that has been driven through the conductor by the pressure of the volts. A milliampere is the thousandth (0.001) part of an ampere.

Watt.—The watt (named in honor of James Watt, a Scotch physicist, 1730–1819) is the unit of electric power. The number of amperes multiplied by the number of volts will give the number of watts. A current of 10 amperes at 100 volts' pressure gives

1.000 watts or 1 kilowatt. A current of 10 amperes at 110 volts' pressure will give the same number of watts as a current of 5 amperes at 220 volts' pressure because the product in both cases is the same; thus 5×220 =1100 watts and 10×110=1100 watts.

Formerly the lighting capacity of electric or other kinds of lights was compared with the lighting capacity of candles, and an electric lamp was labeled as being of a designated candle-power, but electric lamps are now graded according to their watt consumption—*i. e.*, according to the amount of electrical power or energy that they use, and, as previously stated, this is found by multiplying the number of amperes, or, in other words, the amount of current, by the number of volts, *i. e.*, the amount of force necessary to overcome the resistance offered to the passage of the current by the conductor and lighting filament.

The current used on a 16-candle power carbon incandescent light on a 110-volt circuit (*i. e.*, a circuit that offers an amount of resistance that requires 110 volts of electromotive force to overcome) is about ½ ampere and on a 220-volt circuit about ¼ ampere. Such a light, therefore, would now be said to have a 55-watt power since 110÷½=55 and 220÷¼=55. A 25-watt tungsten light will replace a 55-watt carbon incandescent light; *i. e.*, it will give as bright a light without requiring any extra energy.

Electrical power, commonly spoken of as *electricity*, is sold on the basis of what is termed a *kilowatt hour;* by this being meant the amount of electrical energy that would exert a power of 1000 watts for one hour, or one of 100 watts for 10 hours, or of 50 watts for 20 hours, etc.

It is the quantity of current and the amount of force necessary to overcome the degree of resistance offered to the passage of the current that are the principal factors in determining the amount of fuel necessary to use in producing the current.

Electrical Measuring Instruments

Instruments for measuring the magnitude of the current are called *ammeters;* those for measuring the electromotive are called *voltmeters.*

The appliances installed in houses for measuring the amount of electrical power used are called *galvanometers* or *wattmeters;* they are, practically, combined ammeters and voltmeters. There are different types of these, but the essential parts in all kinds are a coil of wire, a magnet, a magnetic needle, and a dial. The main service wires and those from which the house wires are distributed are connected to the coil and all the current entering the house passes through it. As soon as the circuit is closed by turning the switch for a light or other purpose, the current begins to flow, and this induces a magnetic field within the coil. This causes the deflection of the needle, which is suspended in the center of the coil, and thus the pointer on the dial is moved. The greater the number of lamps, etc., in action, the greater will be the magnitude of the current entering; consequently, the stronger will be the magnetic field created, and the greater the deflection of the needle.

Fuses

When an electric current is very strong, and under certain other conditions, the conducting wires of circuits and electrical appliances can become so highly heated that not only may they set fire to any combustible substance in their neighborhood, but, also, the wires themselves may be melted. To obviate these dangers, fuses are used. These are made of metals or metal alloys that melt at a comparatively low temperature and, by melting, break the circuit. This accident, which is often spoken of as *blowing out the fuse*, may be caused by too strong a current, or as the result of placing unusual resistance in the path of a current, as, for example, by attaching an appliance that is wired for a current considerably different, either stronger or weaker, from that of the house attachments, or by getting water between the pieces of contact metal in the socket and screw—water being a very poor conductor of electricity offers much more resistance to a current than metal.

Fuses, especially those used in houses, are encased in fireproof material.

Transformers

A transformer is a special variety of induction coil by means of which the strength of a current can be changed. For example, when buildings a long distance from the power house are supplied with electricity, a current must be sent along the main service wires that would be too strong to introduce into the majority of buildings with safety; therefore a transformer is placed where the service wires for the building or buildings and the main wires connect. Small transformers that can be attached to the electric light sockets in the same way as lamps or other appliances are often used in houses when a weaker current is needed for any purpose as, for instance, running electrical toys. It is important, before using new electrical appliances, to ascertain if they have the same watt power as the lamps. The reason for this was given in the section on fuses.

Rheostat

The rheostat is an appliance placed on some kinds of electrical apparatus to control the amount of current passing to them. This, in the case of heating appliances, means also a control of the amount of heat.

There are different kinds of rheostats, but the variety generally used on household electrical appliances consists of series of coils of wire in a small case. Projecting from the interior of the case is what is called a contact arm. This can be moved from right to left, and *vice versa*. When its knob is opposite the word *low*, it is shutting off part of the current and the temperature of the appliance will be at its lowest; when the knob is opposite the word *full*, all the current is entering the apparatus and its temperature is raised.

Static Electricity

The term static (from *stasis* a Greek word signifying standing) is, as previously stated, applied to electricity that is not flowing along a conductor—for example, that made manifest in Experiment 11 by friction. Static electricity has been defined as *electricity in a state of tension*. It is, however, as stated in the preceding chapter, exactly the same kind of electricity as that of the electric current, which we have been considering; the difference being, not in the electricity, but in the conditions under which the electrification is produced.

The chemical battery and the dynamo, as shown in the preceding pages, are, so long as their circuit is closed, and they are prepared for action, driving a constant stream of electrons along their conductors. Machines such as the Winchester static machine, which is much used in electrotherapy, cause their poles, by friction or other means, to become differently charged—*i. e.*, they cause one pole to assume an electro-negative and the other an electro-positive charge. If the poles are then brought into apposition, or if any substance that will act as a conductor is placed between them, an electrical discharge takes place, and the poles then attain a state of electrical equilibrium, and there is no further discharge until one pole is once more made to acquire a positive and the other a negative charge.

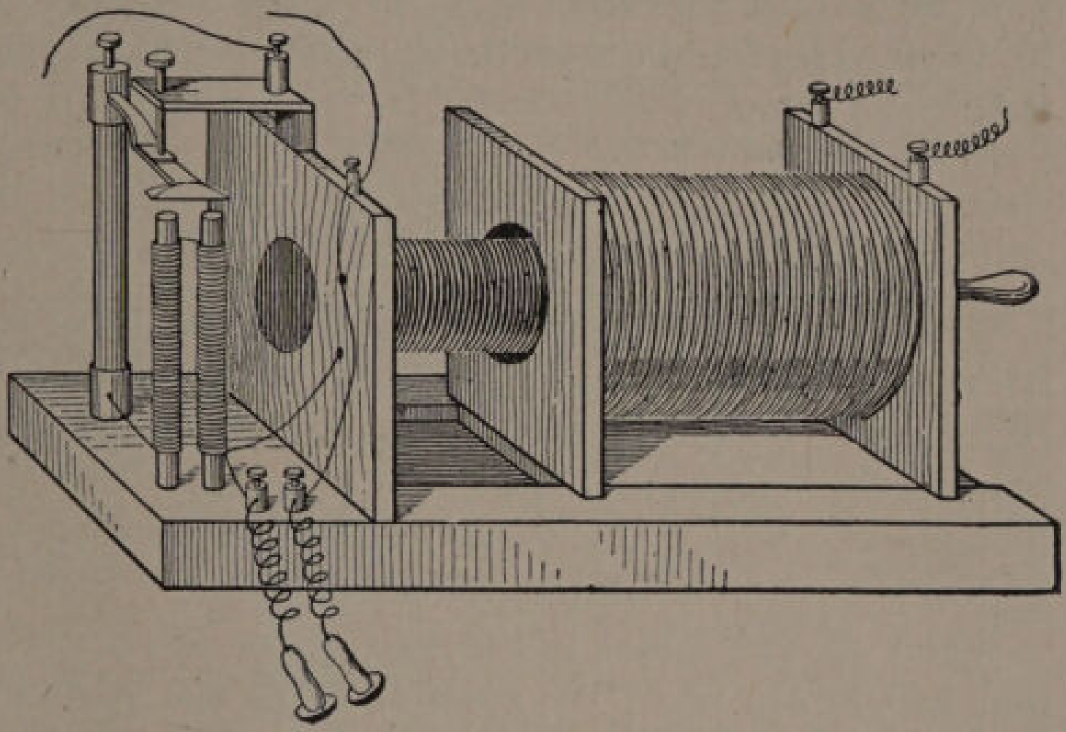

FIG. 53. FARADIC INDUCTION MACHINE.

Thus static electricity is said to manifest itself by attractions and repulsions (this was demonstrated in Experiment 11) and in more or less violent discharges, a discharge bringing about a state of equilibrium.

Nature of lightning.—Lightning is a discharge of electricity between two clouds or between a cloud and the earth, due to one cloud acquiring a positive and the other a negative charge, or, if the discharge is between a cloud and the earth, the difference in charge is between them.

Physiological Action of Electricity

When electricity is applied to the animal body, nerve stimulation results; this induces muscular contraction, which increases metabolism, improves muscle tone, and stimulates the circulation of the blood and lymph. By these means body nutrition and tone may be much improved.

When an individual takes hold of, or is between and in contact with, the electrodes, as the padded poles on the end of the rheophores[1] are called in electrotherapy, he becomes a part of the circuit.

Cathode Rays

When the terminals of an induction coil are sealed into a glass tube from which all air has been expelled, invisible radiations, called *cathode rays*, are found to be emitted by the cathode terminal. These fill the tube with a brilliant yellowish fluorescent light. Cathode rays are not, like light rays, reflected, re-

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[1] The circuit wires of a battery used in electrotherapy are so called.

fracted, or polarized, but they are repelled by negative electric charges, and several experiments have shown them to act in every way as negative electric charges; they are therefore thought to be streams of electrons shot off from the cathode terminal.

X–Rays or Röntgen Rays

In 1895, Conrad Wilhelm Röntgen, of Wurtzburg, Germany, discovered that if the cathode rays struck upon the walls of the vacuum tube in which they were generated or upon a platinum screen or other obstacle placed within the tube, a new kind of rays resulted. These he named *X-rays*, but they are now often spoken of as *Röntgen rays.* X-rays, like cathode rays, are neither reflected, refracted, nor polarized, but, unlike cathode rays, they are not deflected by a magnet nor repelled nor attracted by electric charges; thus, they do not possess the characteristics of either light or of electric charges. It is thought that they are simply irregular pulses or vibrations in the ether caused by the sudden stoppage of the flying cathode rays, but this has not been proved.

X-rays effect the same chemical changes in a photographer's plate as light and hence can be used for taking photographs. A peculiarity of X-ray photographs is that there is only a faint outline of the softer tissues, while hard substances, as bone, stand out clearly. This is because the rays have the power of passing through, not only the glass walls of the tube in which they are generated, but also, though in varying degrees, such substances as the tissues of the body. As hard matter, such as bone and metal, is more or less impervious to these rays, the outline of fractures or of abnormal growth upon the bone will be clearly defined in a photograph, and, if present, any dense, foreign substance, as a needle or bullet, will show. Pus and other exudations in a part will obscure the outline and produce a faint shadow, and thus it is possible to detect abscess formations and the like by means of X-ray photographs. Bismuth is very impervious to these rays, and, if it is swallowed with food, the course of the latter through the stomach and intestines can be followed. This has greatly facilitated the study of the mechanism of digestion and is often of great assistance in diagnosing the nature of gastric and intestinal disturbances.

X-rays are used also for therapeutic purposes in the treatment of such diseases as lupus, cancer, chronic ulcers, and exophthalmic goitre. The exact nature of the action of the rays on the tissues and abnormal deposits, which often results in the alleviation or cure of such conditions, is as yet not well understood, and neither is it known if the X-rays have any direct effect upon micro-organisms in the tissues.

The fluoroscope.—This is a device used for making examinations with the X-rays. It consists of a tapering light-tight box, one end of which is provided with a cuff which will fit closely around the face of the examiner, and the other with a screen of cardboard or similar substance covered with fluorescent matter such as calcium tungstate and barium-platino-cyanide. If a portion of the body is placed between the X-ray tube and the fluoroscope screen, a view of the bones and of any foreign substance, such as would be photographed, will be seen, if present in that part. The need for the fluorescent material is that the X-rays do not themselves affect the optic nerve in a way to give rise to vision, but the light which the rays cause the fluorescent material to emit does.

Radio-Rays

In 1896, Henry Becquerel of Paris found that uranium and certain of its salts and other compounds emitted rays that acted in much the same manner as X-rays, and, in 1898, Madame Curie of Paris found that thorium did likewise, and she found later that a substance contained in pitchblende, in very small quantities, which she named *radium*, possessed the same properties to a very much greater degree. Substances which have this property are said to be *radio-active.*

The rays emitted by radioactive substances are of three kinds, and have been named *alpha*, *beta*, and *gamma rays*.

The alpha rays, which have very little penetrating power, are atoms of the gaseous element helium. The beta rays are thought to be negative electrons and thus they correspond to the cathode rays. The gamma rays, it is thought, are the same as the X-rays; they are produced by the beta rays striking against solid particles.

Salts of radium, chiefly bromides and sulphates, are sometimes used in the treatment of various diseases. The beta and gamma rays penetrate deeply into the tissues; the exact nature of their action is still, however very imperfectly understood.

CHAPTER IX

SOUND—HEARING—SPEECH

Origin of Sound—Some Important Differences between Sound and Light Waves—Transmission of Sound—Why Sounds Are Heard—Causes of Differences in Sounds—Methods of intensifying Sound—Echoes.

Origin of sound.—Sound can be produced by any object that will start wave-like motion in matter that is capable of stimulating the auditory nerves. Objects producing sound are usually in a state of vibration.

Three points of difference between light waves and sound waves.—If an electric bell is suspended in a glass bell jar to which an air pump is attached and the circuit of which the bell forms a part is closed, the strike of the bell hammer will be both seen and heard. If, however, all the air is removed, by means of the air pump, the bell hammer will be seen, but not heard, to strike. This demonstrates one of the great differences between light waves and sound waves—viz., light waves, being vibrations of the ether, will occur in a vacuum, but sound waves, being caused by vibrations in matter, cannot occur in a vacuum.

A second point of difference between light waves and sound waves is the rate at which they travel. Light travels, it will be remembered, at the rate of 186,000 miles a second. The rate at which sound travels varies considerably in different kinds of media, but it takes about five seconds to travel one mile in air. It is for this reason that thunder is heard after the flash of lightning is seen, though, as the lightning is merely the spark occasioned by the discharge of electrons and thunder the waves arising in the air as the result of the vibration occasioned by the discharge, both phenomena originate at the same time. Thus the distance of the electrical disturbance can be estimated by counting the number of seconds that elapse between the flash of light and the sound of the thunder; for, as light travels so quickly, the flash is seen the instant it occurs and, if it is possible to count

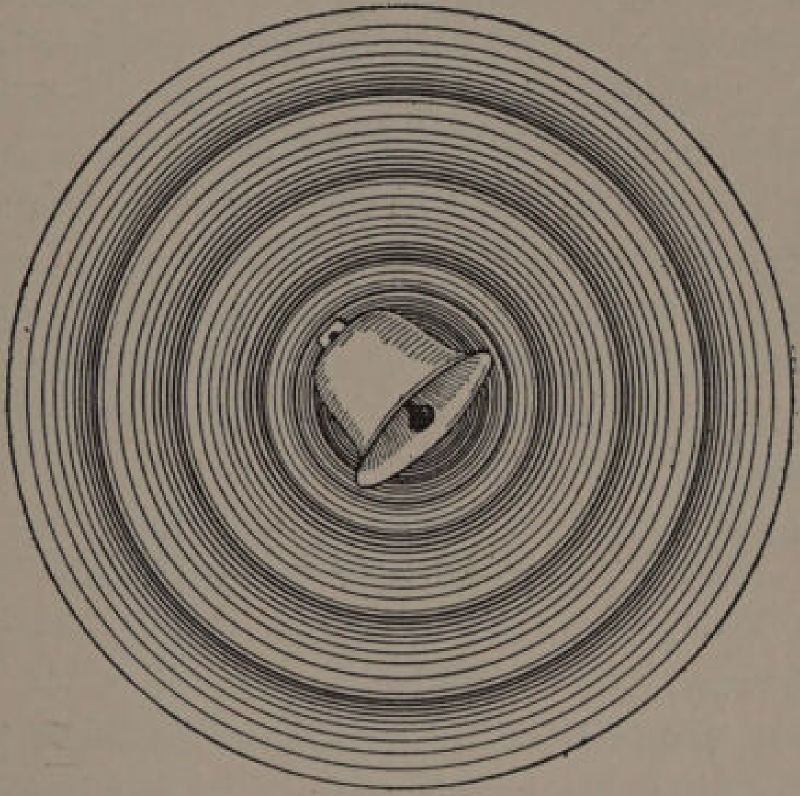

FIG. 54. THE BEGINNING OF A SERIES OF SOUND WAVES STARTED BY A RINGING BELL.

five before the thunder is heard, the electrical discharge must have taken place a mile away.

The third important difference between light and sound waves is the nature of the waves. That of light waves has been already discussed. Sound waves, like light waves, are produced by vibrations, but the vibrating body giving rise to sound causes the molecules of air, or other matter, surrounding it to become compressed; these molecules then rebound and there is more space between them than there was before they were compressed. At the same time, they cause compression of the molecules by which they are surrounded, and these, in turn, act in like manner, and so on; thus, sound waves consist of series of alternate spherical compressions and rarefications. The conduction of sound is analogous to the conduction of heat in as much as the molecules do not move onward, but simply by their own motion produce movement in the molecules by which they are surrounded.

Sound transmitted by matter in its three physical states.—Air is the usual medium for the transference of sound, but it has been found to travel through denser media more rapidly and to a greater distance than through air. For examples: The vibrations of the paddle wheel of a steamer in motion, when the ear is under water, can be heard a mile away; if a long stick of wood with one end held against the ear is scratched with a nail, the sound will be heard by the person whose ear is in contact with the wood, though not by others standing equally near the source of the sound; the sound waves caused by footsteps, etc., will pass a much greater distance in the ground, than in the air, which is why Indians put their ears to the ground when listening for distant sounds; other conditions being similar, distant sounds will be heard more distinctly at the seashore than at high altitudes where the air is less dense.

For sound waves to be transmitted by liquids and solids, however, the sounding body must, as a rule, act directly upon them—for example, sound waves would not induce waves in a stone wall, but would be thrown back from it, as will be seen later.

The mechanism of hearing.—Though sound waves in the air will not themselves induce waves in another medium, they will start bodies that are specially arranged vibrating and thus transmit sound. It is upon this power that hearing depends. For example, when a person speaks, his vocal cords, by their vibratory motion, cause compression of the molecules of the surrounding air and thus give rise to sound waves that pass by conduction, as already described, to the hearer's ears. On entering the ear, the waves impinge upon the drum membrane that is stretched across the opening between the external and the middle ears and start it vibrating. This membrane transmits its motion along the three small bones which stretch across the middle ear and connect with the membrane that separates the middle and the inner portions of the ear. The vibrations of this membrane produce waves in the liquid within that part of the inner ear known as the *cochlea*. These waves cause movement of some of the fine fibrils stretched within the cochlea and thus stimulate the auditory nerves over which impulses are conducted to the center in the brain where sounds are interpreted.

The telephone.—The telephone is an example of a more complicated means of transmitting sound. A telephone, omitting details, may be said to consist of: (1) a box-like portion known as the *transmitter*—this has a mouthpiece in front and a diaphragm at the back and a portion of it is filled with granular pieces of carbon into which the electric circuit wires enter; (2) wires, which are in connection with others from the electric current generator and from the receiver; (3) the receiver, *i. e.*, the part from which the sound is heard. The structure of the receiver is similar to that of the transmitter. When a person speaks into the mouthpiece of the transmitter, the sound waves produced start the diaphragm at the back of the transmitter vibrating; this causes fluctuations in the electric current that correspond with the sound waves produced by the voice, and, though transmitted from one wire to another at the central station, these cause the diaphragm in the receiver to vibrate and give rise to sound waves in the air in the receiver similar to those started in the transmitter by the voice of the speaker.

The phonograph.—The phonograph is still another example of the manner in which sound waves can act upon and be reproduced by different media. Records are made by talking, singing, etc., in front of a cone at the back of which is a membranous diaphragm. The latter is set in motion by the sound waves produced by talking, etc., and transmits its motion to a sharp-pointed instrument, known as the *stylus*, which is attached to its under surface. While the record is being made a wax-coated disc or cylinder is kept rotating under the stylus and the latter cuts into the wax the impressions of the wave forms received from the diaphragm.

The essential parts of phonographs used to reproduce sounds are the same as those on which the records are made, and when a record is placed and made to revolve under the stylus, the latter passes in the traces made by the other stylus and, by so doing, moves the diaphragm, causing it to produce the same form of waves as those which made the tracings on the record and it thus gives rise to similar sounds.

The cones of the phonograph being different in every respect from the human pharynx, there is, however, a varying amount of difference in the tone of sounds reproduced by phonographs.

Difference in sounds.—The differences in sound are classed under three headings—viz., pitch, quality, and loudness.

Pitch.—By pitch is meant *the position of a sound in the musical scale.* The pitch is said to be high when the note is up toward the treble and low when it is down toward the base. The greater the number of vibrations per second, the higher will be the pitch of a sound. The difference between the pitch of male and female voices is due primarily to the different size of the larynx and vocal cords in the two sexes, the larynx being larger and the cords longer in men than in women. A boy's larynx resembles the female larynx and therefore his voice is similar. The so-called *breaking of the boy's voice*, which occurs at puberty, is due to faulty neuro-muscular control brought about in part by the very rapid growth of the larynx and vocal cords which takes place at that time.

Quality, tone or timbre.—The quality, known also as the tone or timbre, of sound constitutes the various characteristics other than pitch by which one person's voice can be distinguished from another's and one type of musical instrument from another. Naturally, this depends largely upon the nature of the part or instrument in which the sound is produced.

Differences in the nature of the material of which musical instruments are made and in the size, shape, and position of their different parts and the tension of their strings will completely alter the nature of the sounds that they can be made to produce.

Differences in voices are due chiefly to differences in the size and shape of the vocal cords, larynx, pharynx, and mouth, and in the degree and kind of control that the individual has of the muscles that enter into the formation of, and surround, these parts. The muscles being pliable and elastic, the shape of the pharynx and other parts concerned can be altered in speaking and singing, and the methods of doing this most effectively constitute one of the most important fundamental principles of voice culture.

Also, differences in the sounds of different letters and words are due to differences in the degree of tension and in the position and movements of the vocal cords and of the muscles of the pharynx, mouth, and tongue during articulation.

Loudness.—Loudness or intensity of sound depends chiefly upon the degree of force exerted by the vibrating body and the nearness of the hearer to the source of the sound. It is also greatly influenced by the density of the medium transmitting the sound (see page 140), the presence of reflectors, and the concentration of sound waves (see page 145).

Noise.—By a noise is meant a loud sound that is disagreeable to the hearer. A noise is usually the result of lack of regularity and rhythm of the vibrations giving rise to the sound.

Reflection of sound.—Sound waves, like light and heat waves, can be reflected or thrown back by objects upon which they strike. Thus, the presence of a sounding board behind a speaker makes it easier to hear him, since the reflected waves serve to reinforce or strengthen the original waves. The walls of a room, if they are not too far from the speaker, act in the same way, and thus it is usually easier to hear in a room than in the open air. If, however, the walls of a room or hall are more than fifty-six feet from the speaker there may be an echo. The reason for this is as follows: The sensation of sound persists for about $1\frac{1}{10}$ of a second, and as sound waves, under ordinary atmospheric conditions of temperature and pressure, travel at the rate of about 1100 feet per second, or 100 feet in the tenth of a second, if the walls are not more than fifty-six feet from the speaker, the waves can reach the walls and be thrown back to the audience before the sensation caused by the original waves has ceased. If, however, the hall is so large that the reflected waves do not reach the audience before the original sensation has ceased and yet is not large enough to prevent reflection, the reflected waves will produce a secondary sensation like unto the first and this is what constitutes an *echo*. The nature of the reflecting surfaces greatly influences the reflection of sound. Sound waves are reflected much more strongly from smooth, hard walls than from those which are carved or hung with draperies. For this reason, draperies are sometimes used to prevent echoes.

Megaphone. Stethoscope. Ear-Trumpet.—These are three examples of appliances used to intensify sounds by concentration of the sound waves and, to some extent, by reflection.

The megaphone, a hollow, cone-shaped appliance, starts the waves produced by anyone speaking into it in one direction and thus keeps them from spreading as widely as usual; this greatly intensifies the sound and increases the distance that it travels in the direction in which it was started.

The stethoscope is an instrument used for distinguishing sounds produced within the body. Stethoscopes vary in their structure, but all kinds have a tube on one end of which is an enlarged portion containing a very delicate diaphragm and on the other end one or two projections that fit into the ears of the listener. The portion containing the diaphragm is placed over the part of the patient's body that is being examined, and the impulses of the sounds produced by the heart, lungs, air in the cavities, or whatever is causing sounds in the part of the body being examined, start similar vibrations in the diaphragm of the stethoscope, and the waves produced in the air in the tube by these vibrations, being confined within a limited area, are so intensified that even slight abnormal differences of sound can be distinguished.

The concentration of sound waves by the ear-trumpet so increases their force that they are able to produce vibration in the membranes and small bones of the ear when, as the result of disease or injury, these parts have become so stiffened that they are not acted upon by less forcible waves.

CHAPTER X

CHEMICAL REACTIONS. VALENCE. RADICALS. CHEMICAL FORMULAS AND EQUATIONS

Nature and Causes of Chemical Reactions—Nature of Valence—Radicals—Chemical Formulæ and Equations.

Chemical reactions.—By a chemical reaction is meant the chemical changes that take place between certain substances when they are brought together.

Agents which promote chemical reactions. Catalyzers.—Certain elements have such a strong affinity for each other that, if brought into contact, they will unite, sometimes even leaving other elements with which they are already in combination to do so. Many other elements, however, especially when already combined, do not unite thus readily, and the help of agents, such as light, heat, electricity, is necessary to promote chemical action. As a rule, solution facilitates—in fact, is often necessary for—chemical reactions, and there are many chemical substances and ferments that, in some unknown way, hasten them, but do not themselves enter into the reaction; such substances are called *catalyzers*, and the action is termed *catalysis*. The action of catalyzers will be demonstrated in Experiment 13.

Reason for chemical reactions.—In the preceding paragraph it was stated that some elements have a very strong affinity for each other, and in Chapter II it was said the chemical affinity meant the electrical attraction between different atoms. Thus, the property and action that were named, before their nature was understood, chemical affinity and chemical union are, in reality, electrical attraction and electrical union, and the quality which causes certain elements to combine more readily than others is thought to depend upon the number and arrangement of the electrons of their atoms which give their ions, see page 109, a stronger or weaker negative or positive electrical charge.

Valence

Nature.—It has been found that each of the elements has the power to hold a definite number of atoms of another element or elements in combination. This capacity is known as valence, and valence has been defined as *that property of an element which determines the number of atoms of another element that it can combine with.* The valence of an element is measured by the number of hydrogen atoms with which one atom of the element combines to form a molecule. Hydrogen and all elements that combine with hydrogen atom for atom are called *monovalent* or *univalent elements;* those which combine with two atoms of univalent elements are called *divalent* or *bivalent;* those which combine with three, *trivalent;* those which combine with four, *tetravalent;* those which combine with five, *quinquivalent.* Some elements, under differing conditions, are able to exert different valences; thus there are such compounds as $FeCl_2$ (ferrous chlorid) and $FeCl_3$ (ferric chlorid). Such elements are said to *exert a variable valence.* In chemical reactions, elements that have the same valence replace each other atom for atom.

In order to explain the nature of valence, it is likened to arms or bonds and the atoms are pictured as having a varying number of arms with which they can hold on to each other. Thus, atoms of the univalent elements, such as hydrogen, sodium, potassium, are considered as having but one arm, each atom of any one of these elements being able to hold on to—combine with—only one atom of another one-armed element; but the oxygen atom, for example, has two arms—is divalent—and can therefore hold on to two univalent or one-armed atoms, but only one divalent; likewise, the atoms of a trivalent element, having three arms, can hold on to three univalent atoms, or one univalent and one divalent, but only one trivalent, and so on.

Valence due to electrical charges.—These arms are of course merely figurative, the real agent controlling the combining of the atoms of, at least, many of the elements[1] is the electrical charge of the ions of the elements. For instance, when a hydrogen-containing compound is decomposed in water or other dissociating liquid—see page 109—the hydrogen will have what is said to be *one positive charge* and, for this reason, it will unite with one atom of an element the

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[1] It has not been proved that this is true of the atoms composing the molecules of organic substances which are not decomposed by the electric current.

ion of which has, what is reckoned as *one negative charge*, or, if the negative charge of the ion is stronger, *i. e.*, if it has two negative charges, two atoms of hydrogen will unite with it. For example, a chlorin ion has one negative charge and it and hydrogen, which has one positive charge, combine thus—HCl, but an oxygen ion has two negative charges and it and hydrogen combine thus—H_2O.

Stable and unstable compounds.—When the atoms of an element are combined with as many other atoms as they can hold on to, the element is said *to be satisfied*, and unless all the elements of a compound are satisfied, the compound is unstable; for example, the carbon atom is tetravalent and, therefore, to be satisfied, it must be combined with four monovalent or two divalent, or one monovalent and one trivalent atoms, or one tetravalent atom. For this reason, carbon dioxid (CO_2) is a stable compound, but carbon monoxid (CO) is unstable and, as soon as it comes in contact with any element that it can unite with, it will do so. For another example, water (H_2O) is a stable compound, but hydrogen peroxid (H_2O_2) is unstable and will very readily become decomposed into water and oxygen.

Radicals.—In many common compounds, two or more of their atoms are very strongly united and do not under ordinary circumstances become dissociated in chemical reactions, but act, in every respect, as though they were one atom. Such a combination of atoms is termed a *radical*.

The valence of some of the common elements, ions and radicals, can be seen in the following table[1]:

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[1] *Elementary Chemistry*, p. 142, Hollis Godfrey. Longmans, Green, & Co.

Element	*Symbol*	*Common Valence*	*Element*	*Symbol*	*Common Valence*
Bromin	Br	1	Zinc	Zn	2
Chlorin	Cl	1	Nitrogen	N	3
Hydrogen	H	1	Carbon	C	4
Iodin	I	1			
Mercury	Mg	1	*Ions and Radicals*		
Potassium	K	1			
Silver	Ag	1	Ammonia radical	NH_3	1
Sodium	Na	1			
Barium	Ba	2	Hydroxyl radical	OH	1
Calcium	Ca	2			
Copper	Cu	2	Carbonate ion	CO_3	2
Iron	Fe	2 & 3			
Lead	Pb	2	Nitrate ion	NO_3	1
Oxygen	O	2	Sulphate ion	SO_4	2
Sulphur	S	2			

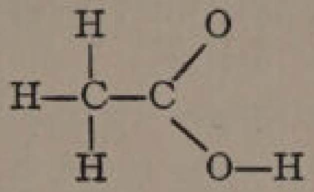
Chemical Formulas

Definition.—By a chemical formula is meant *a combination of symbols used to show the chemical composition of a substance.* In order to read or write a formula correctly, it is necessary to know the symbols of the elements—these were given in Chapter II.—and to understand the meaning of certain signs such as the following: A number placed after a symbol, as stated in Chapter II., shows that in the molecule there are as many atoms of the element as are specified by the figure, but a number placed before two or more symbols shows that there are the number of atoms of all the elements represented by the symbols following the number. For example, Na_2CO_3 shows that there are two atoms of sodium and three of oxygen in sodium carbonate, but $3Na_2CO_3$ implies three molecules of sodium carbonate. Sometimes instead of writing a number before the symbols, as in

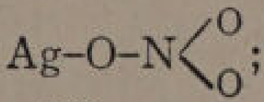
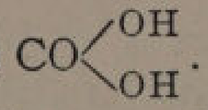
$2Na_2CO_3$, the symbols of groups of elements, of which there are a specified number, are enclosed in parentheses. This method is more especially used when the group of elements constitute a radical or a simple molecule forming part of the molecule of a complete substance. For example, the formula for glycerin may be written $C_3H_5(OH)_3$, there being three OH radicals in a molecule of glycerin; and the formula for the fat known as stearin, may be written $C_3H_5(C_{17}H_{35}CO_2)_3$ showing that in a molecule of stearin there are three molecules, each one of which consists of $C_{17}H_{35}CO_2$, attached to one molecule consisting of C_3H_5. Occasionally a period is placed before the symbol of an element or group of elements, instead of using a parenthesis, and a period is often used to mark off an element or elements that play a special part in a reaction; thus $C_2H_3O_2.H$. The letter n placed after the symbol of an element or elements signifies an unspecified number of atoms, radicals, or molecules, as the case may be; for instance, $C_6H_{10}O_5n$ implies that a molecule of starch consists of a number of molecules each one of which consists of $C_6H_{10}O_5$. When, as in the case of starch, the exact number of simple molecules forming the more complex molecule is not known the letter x is often used, instead of the letter n.

In writing the formulæ of inorganic compounds, the symbol of the electro-positive element, as hydrogen and the metals, is written before that of the electro-negative element or radical, thus, potassium hydroxid is written KOH, and not OHK.

Different kinds of formulæ.—Different groupings of the elements are often made in writing formulæ in order to bring out some special point regarding the structure or action of the molecule. For instance, acetic acid may be written in all of the following ways: $C_2H_4O_2$, this shows simply the elements and the number of atoms constituting one molecule of the acid; or CH_3COOH, this is to show the presence of what is called the *carboxyl radical* which is the characteristic radical of organic acids; or $C_2H_3O.OH$, this serves to show that in acetic acid one hydrogen and oxygen are always linked together; $C_2H_3O_2.H.$ or $H.C_2H_3O_2$, are used to show that in chemical reactions there is one hydrogen atom that is always replaced by another element or radical;



shows the manner in which, it is thought, the atoms are linked to each other in the molecule.

A few other examples of the use of different formulæ to represent the same compound are as follows: Water is written H_2O and $H{-}O{-}H$; nitrate of silver is written $AgNO_3$ and  $Ag{-}O{-}N{<}^{O}_{O}$; carbonic acid is written CO_3H_2 and  $CO{<}^{OH}_{OH}$.

Formulæ which show merely the elements and the number of atoms in a molecule are called *simple formulæ*; those which show some conclusions that chemists have formed regarding the position of atoms in molecules are called *structural* or *constitutional formulæ*. Chemists have formed such conclusions from the way in which molecules act in chemical reactions. Some other examples of structural formulæ will be found on page 159, and the simple formulæ for some common compounds will be found below.

Equations

Purpose and nature.—An equation is used to show the nature of a chemical reaction in a concise form. In the writing of equations, chemical symbols are used. For example: $NAOH+HCl=NaCl+H_2O$ is the equation for the reaction that takes place when sodium hydroxid and hydrochloric acid are put together, and it is a concise and clear way of writing that sodium hydroxid plus hydrochloric acid forms sodium chlorid and water.

Reversible equations.—In many chemical reactions there will be sometimes some action going on in a reverse direction; an equation to show that such action is taking place is written as follows: $AgCl \leftrightarrows Ag+Cl$.

By-Products

In nearly all chemical reaction there are at least two substances formed. The most important one of these is called the *principal product*, and the other or others the *by-product* or *by-products*. In writing an equation, the principal product is usually written first.

Formulæ for Some Common Compounds

Acetic acid. $C_2H_4O_2$.
Alcohol (ethyl). C_2H_6O.
Alcohol (wood). CH_4O.
Aluminium chlorid. $AlCl_3$.

Ammonium carbonate. $(NH_4)_2 CO_3$.

Ammonium chlorid. NH_4CL.

Ammonium hydroxid or ammonia water. NH_4OH.

Barium chlorid. $BaCl_2$.

Calcium chlorid. $CaCl_2$.

Calcium oxid (or lime). CaO.

Calcium hydroxid (or, when in solution, lime water). $Ca(OH)_2$.

Calcium sulphate. $CaSO_4$.

Chlorid of lime (or hypochlorite of lime or bleaching powder). $CaOCl_2$.

Chloroform. $CHCl_3$.

Dextrin. $C_6H_{10}O_5$.

Dextrose. $C_6H_{12}O_6$.

Ether. $(C_2H_5)_2O$.

Ferric chlorid. $FeCl_3$.

Ferrous sulphate. $FeSO_4$.

Formaldehyd. CH_2O.

Glycerin. $C_3H_5(OH)_3$.

Glycogen. $C_6H_{10}O_{5x}$.

Hydrochloric acid. HCl.

Hydrogen peroxid. H_2O_2.

Lactose. $C_{12}H_{22}O_{11}$.

Levulose. $C_6H_{12}O_6$.

Magnesium sulphate. $MgSO_4$.

Mercuric chlorid (or corrosive sublimate or bichlorid of mercury). $HgCl_2$.

Mercurous chlorid or calomel. $HgCl$.

Nitric acid. HNO_3.

Potassium bromid. KBr.

Potassium carbonate (potash, lye). K_2CO_3.

Potassium chlorate. $KClO_3$.

Potassium hydroid. KOH.

Potassium iodid. KI.

Potassium nitrate (saltpeter). KNO_3.
Potassium sulphate. K_2SO_4.
Silver nitrate. $AgNO_3$.
Sodium bicarbonate. $NaHCO_3$.
Sodium carbonate. Na_2CO_3.
Sodium carbonate crystals. Na_2CO_3. $10H_2O$.
Sodium chlorid. $NaCl$.
Starch. $C_6H_{10}O_5$.
Sugar (cane, maple, etc.). $C_{12}H_{22}O_{11}$.
Sulphuric acid. H_2SO_4.

CHAPTER XI

MAIN DIVISIONS OF CHEMISTRY. CARBON. HYDROCARBONS AND THEIR DERIVATIVES

Definitions of Organic and Inorganic Chemistry—Characteristics of Organic Compounds—Isomers—Carbon Oxids—Hydrocarbons and Some of Their Important Derivatives—Alcohols—Aldehyds.

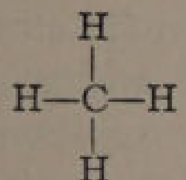
THE study of chemistry has been divided into two main branches, viz.: inorganic chemistry, that dealing with non-living matter such as minerals; and organic chemistry.

Formerly, organic chemistry was said to be *the chemistry concerned with living matter or matter produced as the result of living processes.* In the days when this definition was given, the chemist's work was largely analytical—*i. e.*, the separation of the compounds being studied into their component parts or elements; but now, methods of synthesis—*i. e.*, the building of complex substances from elements or simple compounds by chemical reactions—are very common and every year adds to the number of compounds synthesized in chemical laboratories that it was once thought could be made only as the result of the life processes of living cells. Thus, the long accepted definition for organic chemistry has become incorrect.

A definition that was proposed when the inadequacy of that hitherto used became apparent was *The* *chemistry of the carbon compounds*, because all organic compounds contain carbon, but so also do a few inorganic substances, therefore this definition, though often used, is not very descriptive.

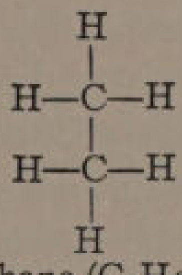
Some important difference between organic and inorganic substances.—Though the supposed difference in matter which gave rise to the division of chemistry into the two branches referred to in the preceding paragraphs, has proved, to some extent at least, to be non-existent, there are marked differences between the substances classed under the two headings. Some of these are as follows: (1) Nearly all inorganic substances are, when in solution, good conductors of electricity, and they are dissociated by the electric current into positive and negative ions, as described in Chapter VII. On the contrary, very few organic compounds are thus affected by electricity. (2) The molecules of inorganic compounds do not contain a large number of atoms, but many of the molecules of organic matter contain a very large number, even thousands, of atoms and the atoms may be combined in many different ways within the molecules. For this reason, there are many organic substances, having different characteristics, that consist of molecules made up of the same number of atoms of the same elements. Such substances, it will be remembered, are called *isomers*. (3) Though there is as much, if not more, variety in organic, as in inorganic compounds, there is exceedingly little variety in the elements composing them, all organic substances being composed of two or more of the following elements: carbon, hydrogen, oxygen, nitrogen, sulphur, phosphorus, and iron. (4) All molecules of organic substances contain carbon directly associated with either hydrogen, oxygen, or another carbon atom, while very few inorganic substances contain carbon, and the carbon in those which do is not combined with the same elements that it is in the organic compounds.

Why it is possible for so many different substances to be formed with the same elements.—One of the chief fundamental reasons for the production of so many different compounds from the same elements is that carbon atoms can combine with each other and in doing so use only one of their valences, thus leaving three links or arms free to combine with other elements; this permits of very large molecules being formed. Also, the relative position of the carbon atoms in the molecule may differ. This difference, the chemist endeavors to show by certain structural formulæ, which, judging by the way in which reactions occur, it has been conceived, may be somewhat the way in which the atoms are linked together in the molecules. The following are examples of such formulæ:



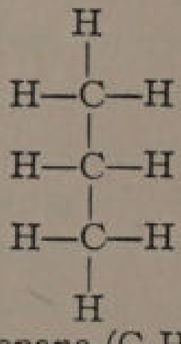
$$\begin{array}{ccccc} & & H & & \\ & & | & & \\ H & - & C & - & H \\ & & | & & \\ & & H & & \end{array}$$

Methane (CH_4)



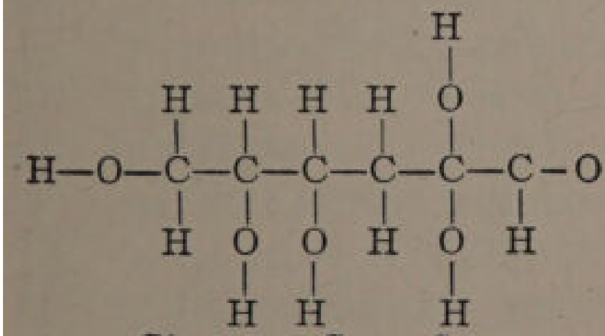
$$\begin{array}{ccccc} & & H & & \\ & & | & & \\ H & - & C & - & H \\ & & | & & \\ H & - & C & - & H \\ & & | & & \\ & & H & & \end{array}$$

Ethane (C_2H_6)



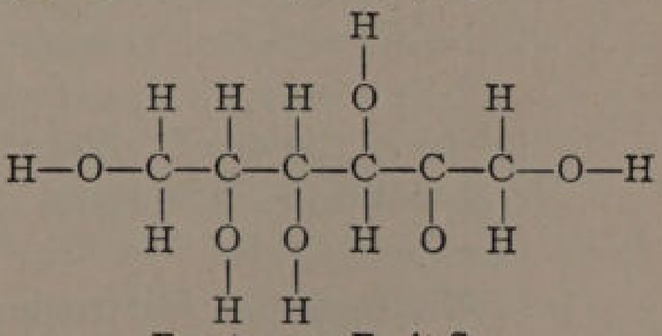
$$\begin{array}{ccccc} & & H & & \\ & & | & & \\ H & - & C & - & H \\ & & | & & \\ H & - & C & - & H \\ & & | & & \\ H & - & C & - & H \\ & & | & & \\ & & H & & \end{array}$$

Propane (C_3H_8)



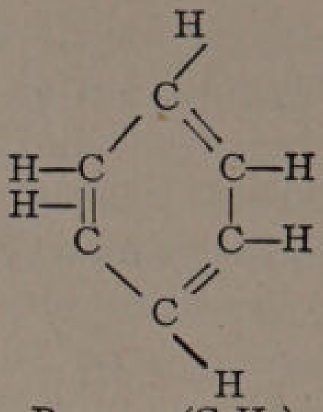
$$\begin{array}{cccccccccccccc} & & & & & & & & & H & & & & \\ & & & & & & & & & | & & & & \\ & H & & H & & H & & H & & O & & & & \\ & | & & | & & | & & | & & | & & & & \\ H-O- & C & - & C & - & C & - & C & - & C & - & C & - & O \\ & | & & | & & | & & | & & | & & | & & \\ & H & & O & & O & & H & & O & & H & & \\ & & & | & & | & & & & | & & & & \\ & & & H & & H & & & & H & & & & \end{array}$$

Glucose or Grape Sugar
($C_6H_{12}O_6$)



$$\begin{array}{ccccccccccccc} & & & & & & & H & & & & & \\ & & & & & & & | & & & & & \\ & H & & H & & H & & O & & & & H & \\ & | & & | & & | & & | & & & & | & \\ H-O- & C & - & C & - & C & - & C & - & C & - & C & -O-H \\ & | & & | & & | & & | & & | & & | & \\ & H & & O & & O & & H & & O & & H & \\ & & & | & & | & & & & & & & \\ & & & H & & H & & & & & & & \end{array}$$

Fructose or Fruit Sugar
($C_6H_{12}O_6$)



Benzene (C_6H_6)

In the first formulæ the C atoms are in chains, in the last one they are in a ring-like form. The *benzene ring formation* is often referred to in chemistry.

Carbon

Carbon is one of the most common elements in nature. It is, as already stated, a constituent of several inorganic and of all organic compounds. It exists in the free state, as well as in combination with other elements; also carbon atoms possess the power of combining with themselves and, though the element carbon is always the same, compounds of carbon atoms can have utterly different appearances; *e. g.*, the diamond consists of nothing but carbon; yet coke, charcoal, and graphite—the black substance in pencils known as lead—are almost, and sometimes quite, pure carbon. Difference in pressure is thought to be an important cause of different combinations of carbon atoms; thus, excessive pressure, it is thought, is one of the chief factors in producing the wonderful combination of carbon atoms that results in the diamond.

Allotropism. Allotropic forms of carbon.—The existence of an element in more than one form with different characteristics is known as *allotropism* and the different forms are said to be *allotropic*. The allotropic forms of carbon are crystalline and amorphous (*i. e.*, without definite shape), the diamond and graphite are crystalline, and coke, charcoal, etc., amorphous substances.

Carbon oxids.—Carbon unites readily with oxygen to form CO_2 (carbon dioxid), a heavy odorless gas. If a carbon compound is burned where there is an insufficient supply of oxygen, CO (carbon monoxid) is formed. The gas burning with a bluish flame, seen between and directly above burning coal, is largely CO.

The Hydrocarbons

All organic substances will undergo disintegration, either in the presence, or more or less complete absence, of air, but both the intermediary and final products of disintegration produced where there is no air differ from those arising in the presence of air. One reason for this is that when disintegration takes place in the air, a large portion of the carbon unites with oxygen, forming CO_2 and, of course, where there is no air, there is no free oxygen to unite with the carbon. Important substances produced in nature by the disintegration of organic substances out of contact with oxygen are the hydrocarbons and coal.

Nature of the hydrocarbons.—The hydrocarbons, of which there are a very large number, consist, as their name implies, of carbon and hydrogen.

Classification.—According to the relative proportion of carbon and hydrogen in their molecules, the hydrocarbons are classified as belonging to the methane, ethylene, and acetylene series. This is shown in the following table:

Methane Series	
CH_4	Methane
C_2H_6	Ethane
C_3H_8	Propane
C_4H_{10}	Butane
C_5H_{12}	Pentane
C_6H_{14}	Hexane
C_7H_{16}	Heptane
C_8H_{18}	Octane

Ethylene Series	
C_2H_4	Ethylene
C_3H_6	Propylene
C_4H_8	Butylene

Benzene Series	
C_6H_6	Benzene
C_7H_8	Toluene
C_8H_{10}	Xylene

Acetylene Series	
C_2H_2	Acetylene
C_3H_4	Allylene

These are only a few of the hydrocarbons known; compounds belonging to the methane series have been found up to $C_{29}H_{60}$.

It will be noticed that the different series differ only in the relative proportion of C and H that the molecules contain, and that the successive substances in each series differ from each other only in the addition of a CH_2.

Physical condition.—In general, at ordinary temperatures, the hydrocarbons with less than 5 atoms of carbon in the molecule are gaseous, those with from 5 to 16 carbon atoms are liquid, and those with a larger number are solid.

How obtained.—Few of the simple hydrocarbons are found free in nature, but they occur chiefly as physical mixtures in petroleum and coal tar and other complex hydrocarbon mixtures, such as gasolene, naphtha, kerosene, natural gas, etc., that themselves are obtained from either petroleum or coal tar by destructive, fractional distillation. That the various constituents of the complex hydrocarbon mixtures can be separated by fractional distillation is due to the fact that they all have different boiling points—the simpler the molecules of the hydrocarbon, the lower its boiling point and thus of the temperature at which it can be distilled, as explained in Chapter IV. Another way in which some hydrocarbons can be prepared is by bringing together carbon and hydrogen compounds that will, under existing conditions, react so that the carbon and hydrogen will combine in the desired proportions; *e. g.*, acetylene gas (C_2H_2) can be made by heating water and calcium carbide together. The chemical reaction that occurs being as follows:

$$\overset{\substack{\text{Calcium}\\\text{carbide}}}{CaC_2} + \overset{\text{Water}}{H_2O} = \overset{\substack{\text{Calcium}\\\text{hydroxid}}}{Ca(OH)_2} + \overset{\text{Acetylene}}{C_2H_2}$$

Source of petroleum and coal tar.—Crude petroleum is obtained in various parts of the world from springs and from wells dug in the earth. Coal tar is obtained from coal by the destructive distillation of coal.

Use of petroleum and coal-tar derivatives.—Many of the hydrocarbons derived from petroleum and coal tar are not only of great value in their uncombined state, but they react with other chemicals, thereby producing a very large number of valuable compounds, some of which are used as medicines, some as disinfectants, others as dyes, and others as a basis for the preparation of various equally useful substances.

Methane and Ethane

Two simple hydrocarbons that form bases for many well-known compounds are methane and ethane.

Methane(CH_4), commonly known as marsh gas and fire damp, occurs free, as well as mixed with other hydrocarbons. It is a colorless, inodorous gas that forms explosive mixtures with air and burns. It is the common cause of explosions in coal mines. It is formed when vegetable matter decomposes where there is little oxygen, and it is found rising from stagnant pools in marshy places where it causes the appearance of bubbles on the surface of the water. It sometimes rises from the earth, together with similar gases, in the neighborhood of petroleum wells, and it is a constituent of natural gas, coal gas, and similar mixtures.

Ethane (C_2H_6) is a gas that rises from the earth with CH_4 and other gases near regions in which there are petroleum springs, and it is found dissolved in crude petroleum.

Methyl chlorid, chloroform, etc.—The forming of these compounds from methane is an example of how this gas, ethane, and other hydrocarbons can be used for the synthesis of more complex substances.

When methane and chlorin are put together in diffused daylight, action takes place gradually in which HCl (hydrochloric acid) is given off and one or more products are obtained according to the length of time the action is allowed to continue. These products are the result of the displacement of one or more hydrogen atoms by chlorin atoms as follows:

$$
\begin{aligned}
&(1)\ CH_4 + Cl_2 = CH_3Cl + HCl \\
&(2)\ CH_3Cl + Cl_2 = CH_2Cl_2 + HCl \\
&(3)\ CH_2Cl_2 + Cl_2 = CHCl_3 + HCl \\
&(4)\ CHCl_3 + Cl_2 = CCl_4 + HCl
\end{aligned}
$$

Methyl chlorid (CH_3Cl) and ethyl chlorid (C_2H_5Cl), which can be made from ethane in the same way as methyl chlorid is made from methane, are highly inflammable, volatile gases that on striking the skin, volatilize so quickly that they freeze the tissue. $CHCl_3$ (chloroform) is the non-inflammable volatile liquid that is much used as a general anæsthetic. Other compounds can be made from methane and ethane by using bromin or iodin, instead of chlorin, as these elements will react with the hydrocarbons in the same manner as chlorin. The yellow antiseptic powder known as iodoform, for example, is CHI_3.

Though methyl and ethyl chlorid, iodoform, and a very large number of similar substances can be thus synthesized in the laboratory, for commercial purposes, they are usually made by using more complex substances that, by chemical reactions, will produce the same combination of elements as those obtained by synthesis. For example, chloroform is made with alcohol and bleaching powder ($CaOCl_2$) and iodoform is made by bringing iodin in contact with an alkali.

Alcohols

Alcohols are sometimes classed with the hydrocarbon derivatives, for they can be synthesized from hydrocarbons, though this is not the manner in which they are made for commercial purposes.

The chemicals necessary for the reaction that produces alcohols from hydrocarbons will not unite with the latter in their natural state and it is necessary to have a preliminary reaction that will form substances which will react with the essential elements. The halogens[1]—either chlorin, bromin, or iodin—are often used for this purpose. For instance:

Methane Chlorin Methyl Chlorid Hydrochloric acid
$CH_4 + Cl_2 = CH_3CL + HCl$
and
Methyl Chlorid Sodium Hydroxid Wood Alcohol Sodium Chlorid
$Ch_3Cl + NaOH = CH_3OH + NaCl$

In the same way ethyl or grain alcohol (C_2H_5OH) can be made from ethane; propane alcohol from propane, and so on.

Methyl or wood alcohol and ethyl or grain alcohol are the alcohols in common use. Ordinarily, the former is made by the destructive distillation of wood (as described in Chapter IV.), and grain alcohol, by the fermentation of glucose, as described in Chapter XIX.

Primary, secondary, and tertiary alcohols.—Alcohols such as those described in the preceding paragraphs contain only one OH radical and are called *primary alcohols*, but alcohols containing two or more OH radicals can be made from the higher hydrocarbons. Alcohols with two OH radicals are called *secondary alcohols*, and those with three, *tertiary*. The only one of such alcohols that need be referred to here is glycerin or, as it is often termed in chemistry, *glycerol*. Glycerin can be made from propane and it has three OH radicals, it is therefore classed as a *tertiary propyl alcohol*. The chemical formula of a primary propyl alcohol is C_3H_7OH and of a tertiary propyl alcohol (glycerin) $C_3H_5(OH)_3$.

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[1] The term halogen is derived from a Greek word meaning the sea. Chlorin, iodin, bromin, and fluorin were so classed because obtained from sea salt or seaweed.

Though glycerin can be made from propane, it is ordinarily prepared from fat, of which, as will be seen in Chapter XVI., it is the base.

Aldehydes

When a primary alcohol is oxidized (see page 174) it is converted first into an aldehyde, then into an acid, and finally into carbon dioxid and water. The difference in the composition of an alcohol and an aldehyde consists of a lack of two H atoms in the latter, since the O, in the process of oxidation, combines with two H atoms of the alcohol to form water (H_2O). Thus:

Alcohols Aldehydes

$$CH_3OH + O = CH_2O + H_2O$$
$$C_2H_5OH + O = C_2H_4O + H_2O$$

The aldehyde formed by the oxidation of methyl alcohol is the disinfectant gas known as *formaldehyde*, and the aldehyde of ethyl alcohol is called *acetic aldehyde*. The acid produced from formaldehyde is called *formic acid*, and that from acetic aldehyde, *acetic acid*. The chemical reactions that occur in the oxidation of these aldehydes and acids can be seen by the following equations:

Aldehydes Acids

$$CH_2O + O = CH_2O_2 + O = H_2O + CO_2$$
$$C_2H_4O + O = C_2H_4O_2 + 2O = 2H_2O + CO_2$$

The aldehydes can be made from alcohol by burning the latter in specially constructed lamps, or by bringing the alcohol in contact with an oxidizing mixture (one that will readily part with some of its oxygen), as potassium bichromate and dilute sulphuric acid.

Formaldehyde gas is soluble in water up to 40%. The solution of the gas is known as formalin. The gas can be quickly liberated from solution by the use of certain chemicals, such as potassium permanganate, which, when added to the formalin, oxidizes some of the formaldehyde and in so doing produces heat enough to volatilize the rest of it. It is often procured in this way for the disinfection of rooms.

CHAPTER XII

OXYGEN

The Occurrence and Nature of Oxygen—Nature of Oxidation, Spontaneous Combustion, Oxids, Fireproof Material, Fire Extinguishers, Products of Oxidation.

Occurrence.—Oxygen is the most common element in nature. It exists free (*i.e.*, not in chemical combination with other elements) in the air; it, in combination with hydrogen, forms water, and, united with different elements, it constitutes nearly one half of the rocks and other substances forming the earth's crust; also, it is one of the essential constituents of animal and vegetable matter.

How obtained.—When the majority of compounds and elements are heated they will take oxygen from the air very readily, but, though oxygen exists free in the atmosphere; it is not easily extracted from it for laboratory purposes or when a large supply of the pure gas is required. Therefore, for such purposes, oxygen is usually obtained by extracting it from some substance that contains a comparatively large amount of it and parts with it readily. The substances most frequently used are potassium chlorate, manganese dioxid, and mercuric chlorid. When potassium chlorate and manganese dioxid are mixed, the oxygen is liberated at a much lower temperature than it is when either of these two chemicals is used alone, and when it is used with the potassium no change occurs in the manganese. By what means the manganese hastens the liberation of oxygen from the potassium chlorate is not known. When used in this way the manganese is called a *catalyzer.* (See page 147.)

Experiment 13. Objects: To study (1) a method of reducing a compound to its constituent parts; (2) a method of liberating a gaseous element from a compound and collecting the gas; (3) the properties of oxygen; (4) the use of a catalyzer.

Articles required: An iron stand with a clasp, a large-sized, hard-glass test tube, a Bunsen burner, a wide-mouthed bottle filled with water, a similar bottle empty, two glass squares large enough to cover and extend slightly beyond the mouth of each bottle, glass tubing bent as in Fig. 55, a pan about ¾ full of water, potassium chlorate (K_2ClO_3), manganese dioxid (MgO_2).

Directions: Some members of the class put about 10 grams of potassium chlorate into their test tubes, and the others 10 grams of potassium chlorate and 3 grams of manganese dioxid (mix the two chemicals on paper before putting them into the test tube). Connect the apparatus as shown in Fig 55, cover the bottle filled with water with a glass square and invert it in the pan of water, being careful not to spill the water while doing so (because of the atmospheric pressure on the water in the pan, that in the bottle will not run out after the bottle is inverted). Apply gentle heat to the part of the test tube containing the chemical. Notice, and record, the time required to liberate the oxygen. When the gas begins to be evolved, insert the free end of the tubing under the bottle; to do so, tip one end of the bottle slightly, being careful not to raise the opening above the surface of the water in the pan. After the gas has displaced all

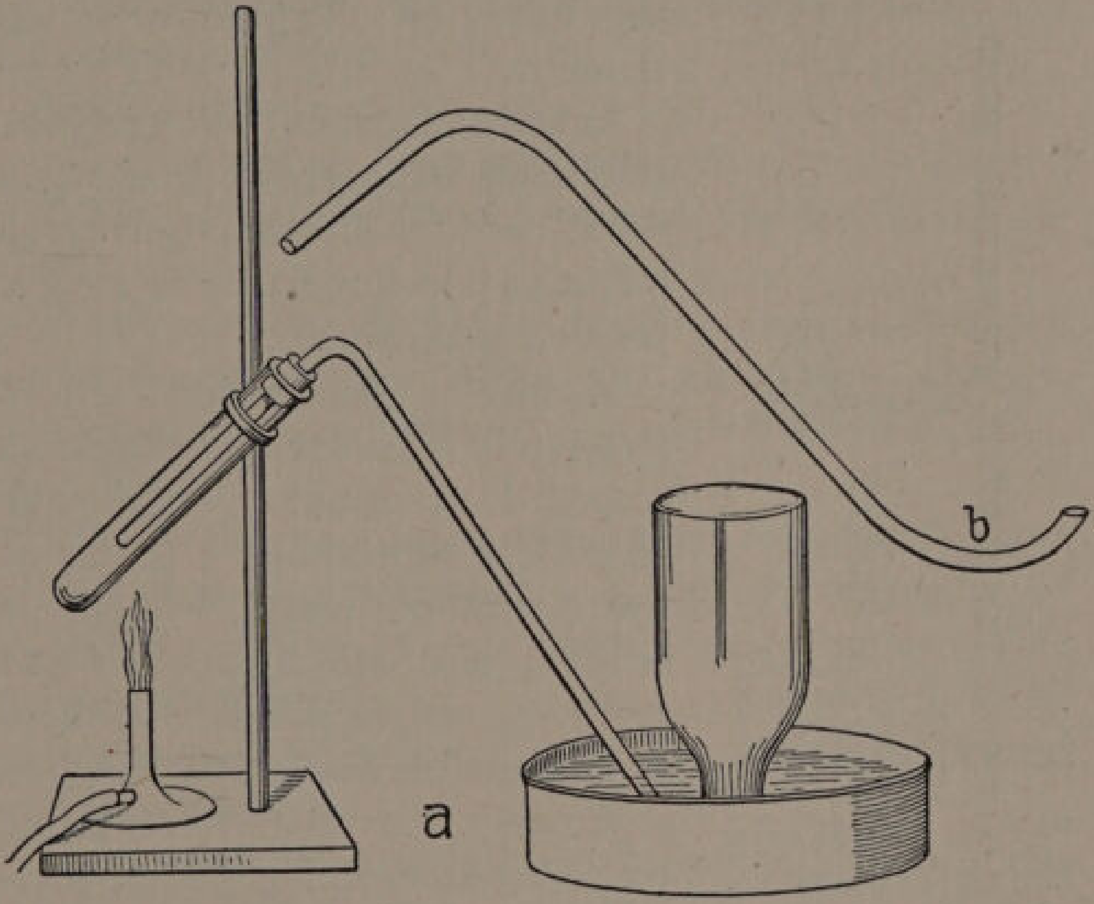

FIG. 55. (a) Apparatus for the preparation of oxygen. (b) Shape of tube.

the water in the bottle, turn the bottle and stand it on the table, keeping it tightly covered with the glass square. Drop a lighted match or, better, a burning strip of wood into the bottle without removing the glass cover. Notice the length of time that the wood burns and the nature of the flame. Drop a burning piece of wood into the empty bottle and cover it at once with a glass square. Notice the difference in the nature of the flame in this bottle and in that with the oxygen and in the length of time that the wood continues to burn. Notice in both bottles if water is formed, if there is any smoke, and the charred condition of the wood.

Experiment 14. Object: To study the result of the union of oxygen with an element.

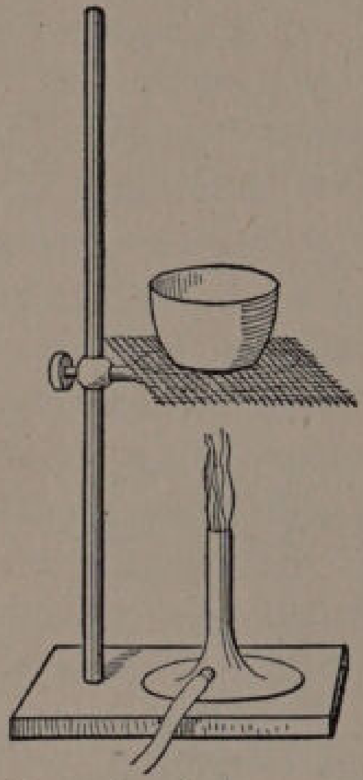

FIG. 56. APPARATUS FOR EXPERIMENT 14.

Articles required: A crucible, a triangle support, an iron stand with ring, a Bunsen burner, a balance, powdered zinc or iron, a pointed glass rod.

Method: (All weights must be absolutely accurate and should be recorded at once.) Weigh the crucible, weigh about 3 grams of the powdered metal, weigh the crucible and metal together and see if the weight tallies with those of the crucible and metal weighed separately. Arrange the crucible on the stand as in Fig. 56, supporting it in the triangle; light the burner and secure a low flame; put this under, but at a distance from, the crucible. If the zinc begins to glow or give off a white smoke, lower the flame instantly, or there will be a loss of substance. If a crust forms on the zinc, pierce it with the pointed glass rod in several places so as to expose a fresh surface to the air; be very careful not to lose any of the powder in doing this, and if particles stick to the rod, tap the latter gently against the crucible so that they will fall back. Continue the heating for about ten minutes and then let the crucible and its contents cool. When the crucible is cool enough to touch, weigh it with its contents. If the experiment has been properly performed there will be found to be a slight gain in weight due to the oxygen which has united with the metal. See under Oxids, page 179.

From the results of these experiments answer the following questions:

Does oxygen burn?

Does oxygen support combustion?

What became of the oxygen in the bottle during the burning?

What was the difference in time required for the liberation of oxygen when the potassium chlorate was used alone and when it was used with the manganese dioxid?

What is the manganese called when used in this way?

Compare what happened to the wood—a compound —with the effect of the union of oxygen with the metal—an element.

Has oxygen color?

Has it odor?

Is oxygen heavier or lighter than air? Pure air consists principally of about 21 per cent. oxygen, 79 per cent. nitrogen, and 0.04 per cent. carbon dioxid. The atomic weights of these elements will be found on page 31.

Is oxygen lighter or heavier than pure water? (Water consists of hydrogen and oxygen.)

Physical and chemical properties of oxygen.—Oxygen is a colorless, odorless, tasteless gas. It is a little heavier than air and than pure water and it is slightly soluble in the latter.[1] It can be liquefied

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[1] Fish will not live in boiled water since during boiling the free oxygen, which the fish need, is expelled.

by putting it under great pressure at a low temperature. At low temperatures, oxygen is rather inert chemically, for it does not combine wth many substances, but at high temperatures, it is exceedingly active. Certain catalyzers, however, as, for example, the enzymes present in both animal and vegetable tissues, permit of oxidation at comparatively low temperatures.

Oxidation and combustion.—The process of the union of oxygen with matter, either compounds or elements, is spoken of as oxidation. Oxidation is always accompanied by the generation of heat. If oxidation takes place very quickly, heat will be evolved so rapidly that the matter may be brought to a white heat (see page 86) or burst into flame. When this occurs, the process is often termed *combustion*. When the union of oxygen with matter goes on slowly, the heat may be developed so slowly that it will be difficult to detect it. The rusting of iron and other metal is an example of this slow oxidation, and iron rust is a form of iron oxid. Other oxidation processes are (1) the oxidation in the body tissues of the substances derived from food; (2) the changing of wine, cider, etc., to vinegar and of vinegar and other organic acids to carbon dioxid and water; (3) some putrefactive processes; (4) the hardening of paint, as will be seen later, is partly due to oxidation.

Kindling temperature.—As stated in a preceding paragraph, oxygen does not ordinarily unite with many substances at a low temperature, and the temperature at which it will combine varies very greatly with different substances, but is always the same for each kind of substance. The temperature at which a substance will begin to combine sufficiently rapidly to *take fire* is called the *kindling temperature* of that substance or, especially in the case of liquids as the hydrocarbons, the *flash point*.

Matches.—The various constituents of matches furnish a good example of differences in the kindling temperature of different substances. Some match heads consist principally of yellow phosphorus and sulphur, and the kindling temperature of phosphorus is so low that even a slight amount of friction will cause it to ignite; this provides sufficient heat to ignite the sulphur, the kindling temperature of which is also low, though not as low as the phosphorus, and the burning sulphur heats the wood to its kindling temperature. It would be too dangerous to have the match head composed entirely of a substance that could be ignited by friction. Even with sulphur, the use of yellow phosphorus for match heads is dangerous, not only because it ignites so easily, but also because it is excessively poisonous, and, therefore, red phosphorus is now often used instead of yellow. Red phosphorus, which is made by heating yellow phosphorus to about 250° C., is not poisonous and it will be ignited by friction only when it is in contact with substances rich in oxygen, such as chlorate and chromate of potassium and peroxid of lead. The head of one kind of *safety match* contains the oxygen compounds and the red phosphorus is on the outer surface of the match box; thus the ignition of a match is impossible, unless it is struck on the box.

Spontaneous combustion.—By this is meant the *sudden ignition of matter at comparatively low temperatures when heat has not been intentionally applied*. For example, hydrocarbons, such as kerosene, have a low flash point, especially when they contain certain impurities that are often present in cheaper grades, and fires have been caused in tenements in the summer time because the heat of the rooms became raised to the degree at which the kerosene combined rapidly with oxygen. Painters' rags wet with oil and turpentine, heaped together, are the cause of many fires, because oil and turpentine having a strong affinity for oxygen keep uniting with it slowly and the rags, being in heaps, prevent the escape of the heat generated by the oxidation, so that the kindling temperature of the oil, which is low, is soon reached and the mass suddenly bursts into flame.

Difference and similarity of the results of the oxidation of compounds and of elements.—As demonstrated by Experiments 13 and 14, when oxygen unites with a compound, the latter becomes dissociated into simpler compounds and its elemental constituents, but when oxygen combines with an element, a compound substance is formed. Judging from appearances, it might be thought that another point of difference was an increase of bulk and weight of the oxidized element and a loss of substance in the compound, but it has been found that when all the gases which escape from a compound as it burns are caught and weighed, as well as the ashes which remain, there is not only no loss of weight in the matter burned, but, just as in the case of the oxidized element, a gain in weight that is equal to the amount of oxygen that has united with the matter in the course of combustion.

Products of combustion.—The products of combustion, of course, vary according to the nature of the matter burned, but, as already stated, all oxidation gives rise to heat and all organic matter, whether of animal or vegetable origin, yields water and carbon dioxid, for all organic substances contain carbon, hydrogen, and oxygen. Also, there will be various other gases combustible and incombustible, which will pass off from the burning matter, and, usually, there will be some uncombined carbon, mineral matter, and other incombustible solids that will, for the most part, remain as ashes, but small particles will be carried off with the gases.

Smoke.—The hot gases, water, and other matter arising from burning substances gradually cool as they pass into the air and in doing so condense to some extent, forming the cloud-like matter known as smoke. When a large amount of incombustible matter and products of incomplete combustion are present, the smoke will be of a dark color and, especially when the burning material is of an oily nature, there will be a considerable amount of matter that will condense easily and in doing so absorb some of the carbon and other solid substances present in the smoke; this constitutes *soot*. The substances contained in soot are not all fully oxidized and, for this reason, if a large quantity is allowed to collect in a chimney, it may take fire, when a fire is lighted in the stove, furnace, or fireplace.

Experiments to study the nature of combustion.—

Experiment 15.

Articles required: A hard-glass test tube, a test-tube holder, two or three small pieces of wood, a cork with a hole in the center, a piece of glass tubing about three inches long drawn to a point at one end, a Bunsen burner, matches.

Procedure: Put the pieces of wood into the test tube, insert the tubing in the cork, leaving the pointed end free, place the cork in the tube and hold the latter so that the wood will be over the lighted burner. When smoke begins to issue from the tubing, apply a lighted match to it. There being combustible gases in the smoke it will ignite. When the gases cease to burn and it is impossible to relight them, remove the charred wood from the test tube, examine it, and hold a piece of it with the forceps in the flame.

Does it burn with or without a flame? See page 179.

Of what does this residue of the wood consist?

Experiment 16.

Articles required: Substitute a few pieces of coal for the wood, otherwise have the same articles as for Experiment 15.

Procedure: Same as Experiment 15.

Experiment 17.

Articles required: A hard-glass test tube, a test-tube holder, about ten grams of cane sugar, a Bunsen burner, matches.

Procedure: Put the sugar into the test tube and hold the latter in the flame. Notice the changes that occur in the sugar. After the sugar has turned brown and smoke issues from the tube, apply a lighted match to the smoke.

How do the products of the burning and burnt coal and wood compare with those of the sugar?

Experiment 18.

Articles required: A wide-mouthed bottle, a glass square, a taper or strip of wood, lime water, a beaker, a piece of glass tubing about four or five inches in length.

Procedure: Let the taper or wood burn in the bottle for some time, keeping the bottle covered with the glass square as much as possible without putting out the light. After a few minutes, put out the light and pour a little lime water into the bottle.

Put some lime water into the beaker and blow into it through the glass tubing.

The white substance that appears in the lime water is calcium carbonate ($CaCO_3$), formed by the union of carbon dioxid (CO_2) and the lime, which is calcium oxid (CaO), in the water. Thus, $CaO+CO_2=CaCO_3$.

Why do you get the same results from the burning of a taper or wood and from the breath?

Flame and incandescence.—Matter may burn with or without a flame. If the matter is a volatile substance, such as ether, kerosene, or alcohol, or if, though not itself volatile, it gives off combustible gases when heated, as do coal and wood, there will be a flame during combustion, but matter such as solid carbon will combine with oxygen and thereby become so hot that it will glow, and metals and oxids may be so highly heated that they will glow—*i. e.*, they will become incandescent—but in neither case will there be a flame. As a rule, flame occurs only in the burning of volatile substances.

Oxids.—An oxid is a compound of oxygen and an element or radical. Thus water (H_2O) is an oxid of hydrogen, and carbon monoxid[1] (CO) and carbon dioxid (CO_2) are oxids of carbon, sulphur dioxid (SO_2) is an oxid of sulphur, and so on. There are only two elements—bromin and fluorin—that will not, under any circumstance, unite with oxygen.

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[1] Oxids which have but one atom of oxygen in the molecule are called *monoxids*, and those which have two atoms are called *dioxids*.

Nature of fireproof buildings, clothing, etc.—Buildings are said to be *fireproof* when they are constructed of substances that are non-combustible, or that will not burn except at excessively high temperatures, or in the presence of free oxygen (such are iron, concrete, stone), and of matter saturated with oxids which have taken up all the oxygen that they can combine with and therefore cannot burn, burning being nothing but the union of matter with oxygen. Wood, clothing, and other absorbent material can be made *fireproof* by being treated with such oxids.

Fire extinguishers.—Since burning consists of the combining of matter and oxygen, if air is excluded from a substance, it cannot burn. Therefore, the quickest way to check a fire is to prevent the access of air to it. An effectual way of doing this is by the use of a fire extinguisher containing chemicals which, when combined, form carbon dioxid. As this gas holds all the oxygen that it can unite with and is heavier than air, it will, if poured over burning matter, act as a non-inflammable, air-proof covering. Many extinguishers contain a flask of sulphuric acid surrounded with sodium carbonate and when the extinguisher is tilted the acid comes in contact with the soda and carbon dioxid is evolved. Thus: $H_2SO_4 + Na_2CO_3 = Na_2SO_4 + H_2O + CO_2$.

CHAPTER XIII

FUELS AND ILLUMINANTS

The Nature and Origin of the Substances Commonly Used for Fuel and Lighting.

Classification of fuels.—Fuels are usually classified as solid fuels, liquid fuels, and gaseous fuels.

The solid fuels are coal, coke, peat, wood, and charcoal.

The liquid fuels in common use are alcohol, gasoline, kerosene or coal oil, and crude fuel oil.

The common gaseous fuels are coal gas, natural gas, gasoline or air gas, and acetylene gas.

Solid Fuels

Classification of coal.—The several varieties of coal are usually classified under three headings, viz., anthracite or hard coal, bituminous or soft coal, and lignite.

Reason for differences in coal.—The differences in coal are thought to be due to the stage of formation in which the coal is when it is taken from the earth.

Origin of coal.—Coal is derived from vegetable matter that through many thousands of years has been undergoing decomposition out of contact with the air and, during the process, losing its hydrogen and oxygen.

Lignite.—Lignite, a brown coal, represents a comparatively early stage in the process of coal formation. The word *lignite* is from the Latin *lignum* meaning *wood*, and it is so called because it still shows traces of wood structure. It contains proportionately less carbon and more oxygen and hydrogen than the other coals, and it therefore gives less heat than they do, because the heat value of fuels depends principally upon the union of the carbon in the fuel and the oxygen of the air.

Bituminous or soft coal.—This coal has passed farther in the stage of decomposition than lignite and has therefore lost more of its ingredients other than carbon. The disadvantages of soft coal for heating purposes are: (1) that being soft and easily powdered, it is very dirty; (2) many varieties cake as they burn, and the caked masses require to be broken constantly in order to allow a free draft through the fire; and (3) there is usually a great deal of smoke and soot from a soft-coal fire. Bituminous coal yields more heat than lignite, but less than anthracite. Medium grades of soft coal are considered the best coal for the manufacture of coke and of coal gas.

Anthracite or hard coal.—In its purest state, this coal has lost nearly all its ingredients other than carbon. Therefore it gives more heat than either of the other two varieties. It burns with little or no smoke, and the best varieties do not produce any clinkers. Sulphur, a common impurity in all kinds of coal is one of the principal causes of clinkers.

Coke.—Coke is produced by the destructive distillation of bituminous coal (see distillation, page 64).

It consists of free carbon mixed with varying amounts of ash. It produces intense heat, but no flame or, if the fire is deep, a carbon monoxid flame. See page 161.

Peat.—Peat results from the decomposition of mosses and similar plants under water. In its natural state, it has not a high fuel value, for it contains a large per cent. of water, but, of recent years, ways of drying it have been devised that have increased its fuel value very considerably.

Classification of woods.—Woods are classed as soft woods and hard woods. Wood, as used for fuel, consists principally of cellulose $(C_6 H_{10} O_{5x})$, mineral matter, and water, and the conifer woods contain a varying amount of resin. If there is much water, the heat value of the wood is lessened because much of the heat will be used in vaporizing the water; see page 45. The wood from trees cut in winter has less water than that from those cut at other times because, in cold weather, there is less sap flowing through the trees.

Soft woods.—The soft woods most used for fuel are pine, cedar, spruce, hemlock, poplar, willow, and redwood. Soft woods ignite easily and are therefore valuable for kindling, but they burn quickly and thus, if used alone, will yield a very hot fire for a short time, but it will require a large amount of wood to keep up the fire. The conifers yield more heat than other soft woods because of their resin.

Hard woods.—The hard woods most commonly used for fuel are ash, beech, birch, sugar and black maple, oak. A cord of hard wood and a ton of coal are said to have about the same heat value. Weight for weight, if properly dried, hard and soft woods yield about the same amount of heat, but hard woods burn more slowly and thus last longer than soft woods.

The mineral matter left after wood is burned consists largely of potassium, phosphorus, and calcium compounds; thus, if the ashes are kept dry and put into the ground they act as fertillizers, since plants need these salts for their growth. If the ashes are wet, however, their value for this purpose is lessened, for the potassium salts, which are the most important ones, are changed to potassium hydroxid.

Charcoal.—Charcoal is obtained from wood in the same manner as coke is from coal and its characteristics are very similar to those of coke.

Liquid Fuels

The liquid fuels in most common use are alcohols, gasoline, kerosene, and other so called *fuel oils.*[1]

Alcohols.—Wood alcohol and denatured ethyl alcohol are the alcohols used for fuel. Wood alcohol, as stated in the preceding chapter, is made by the destructive distillation of wood. Denatured alcohol is ethyl alcohol plus some substance that has been added to it in order to render it unfit for drinking. This is used because it is cheaper than the pure alcohol, since it is tax free, while there is a heavy tax on the latter. Wood alcohol does not give as much heat as ethyl alcohol and it is very poisonous. A new proprietary preparation called *sterno*, in which the alcohol is incorporated in a waxy substance, is now much used instead of the liquid alcohol. Alcohol, in any form, would be an expensive and unsatisfactory fuel

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[1] Kerosene and similar hydrocarbon derivatives, though of an oil-like nature, are not true oils.

to use for large fires, but when only a small flame is required, as for chafing-dish cooking, it is very convenient to use. It is always important that the wick of an alcohol lamp or stove should fit properly, for if it is too small the flame may reach the alcohol and, unless the utensil is constructed for the burning of the alcohol directly, the results may be disastrous.

Gasoline.—This is obtained from petroleum by fractional distillation. Gasoline has a lower specific gravity and is more volatile than kerosene. The latter quality makes it a more dangerous fuel, because it volatilizes at a low temperature and the vapor, when mixed with a certain amount of air, becomes explosive. Quite a large amount of air is required, however, and if the following precautions are observed there is very little more danger attending the use of gasoline than of any other liquid fuel: (1) The stove must be properly cleaned and it must be always filled before it is used; (2) the stove must be in good condition so that there will be no leakage and the wicks must fit the burners; (3) the cover of the tube through which the stove is filled must not be left off, nor even loose; (4) a good quality of gasoline must be used; (5) the light must be put out by turning down the wick, never by blowing.

It is when stoves that have been partially empty for hours are first lighted that trouble usually occurs; because, while the stove has not been in use, the gasoline has been volatilizing and, as the vapor has not been burned, it has been filling the stove and mixing with the air in it. As a high temperature accelerates volatilization, explosions from this cause are most likely to occur in hot weather or in a hot room.

Qualities of good gasoline.—Good gasoline has a neutral reaction, a specific gravity of about .68 to .72, and, when filtered through chamois, it leaves no residue.

Kerosene.—Though kerosene is not as explosive as gasoline, the same precautions are necessary in its use and for the same reasons. It is to be remembered that explosions occur only when the reservoir of a stove or lamp contains definite proportions of air and vapor, or when the hydrocarbon becomes ignited, which it can hardly do if the precautions mentioned under gasoline are taken. Kerosene that has a flash point below 100° F. is not safe for general use, and in many cities where the temperature is likely to be high at times, the law prohibits the use of a kerosene that has a flash point below 125° F. or 150° F.

Kerosene stoves.—The blue-flame kerosene stoves are usually considered the best. In the majority of these, the oil is fed into a hollow ring at the bottom of the burner, where it is heated to a sufficiently high temperature to vaporize it. The vapor becomes mixed with air in the burner, and thus just as was the case with the gas in the Bunsen burner, it gives a blue flame.

Especially in oil-producing regions, crude fuel oils that are sold under various trade names are often used for fuel in specially constructed ranges, but only the more refined hydrocarbons should be used in the ordinary kerosene and gasoline stoves.

Gaseous Fuels and Illuminants

The gases generally used for fuel and lighting purposes are coal gas, water gas, natural gas (the two last named are used for lighting only in connection with some of the higher hydrocarbons), acetylene gas, and gasoline or air gas.

Coal gas.—As already stated, coal gas is a product of the destructive distillation of coal. Its principal constituents are hydrogen, methane, small quantities of other simple hydrocarbons, carbon monoxid, carbon dioxid, nitrogen, and a small quantity of bisulphid of carbon.

Incandescent gas lights.—The nature of the coal-gas flame was discussed in connection with the Bunsen burner, page 9. Also, it was there shown that the presence of free carbon in the flame increased the light-giving properties of the gas, because the carbon could be raised to a white heat and thereby be rendered luminous without being disintegrated or volatilized. As stated in the preceding chapter, matter rendered luminous in this way is said to be *incandescent*. To find a means of increasing the amount of matter that could be made incandescent in a flame was the object of many experiments, of which two well-known results are the *Argand burner* and the *Welsbach mantle*.

The Argand burner—called after the inventor—gives a round flame in which a large amount of carbon is liberated.

The Welsbach mantle—invented by Auer von Welsbach—is made by cutting and sewing some mercerized cotton into the required shape, using asbestos thread for the sewing, and then saturating the cotton with a solution of certain oxids, more especially those of two rare elements known as *thorium* and *cerium*. In the process of manufacture, the cotton is burned out and the liquid in which the oxids were dissolved is evaporated and the oxids thus hardened so that the mantle, when completed, consists of a thin, filmy cap-like structure of oxids. This is connected with a burner that gives a non-luminous, and, consequently, the hottest flame, and this heats the oxids to such a degree that they give a white luminous light six to eight times more powerful than that produced when an ordinary illuminating gas burner is used.

Surface combustion.—Somewhat the same principle as that upon which the structure of the Welsbach mantles was based is used in the *surface-combustion stoves*, though in the stoves heat, and not light, is the thing desired. The stove burners are placed in the midst of incombustible material, such as alumina (Al_2O_3) and silica (SiO_2), and are so constructed that just enough air to allow of the combustion of the gas can enter. The gas burns without a flame and heats the alumina, etc., to incandescence, and the heat reflected greatly exceeds that from an ordinary gas flame.

Natural gas.—This is obtained in certain localities by boring wells in the ground. Methane (CH_4) is its chief constituent, but it contains small quantities of ethane (C_2H_6) and other simple hydrocarbons.

Water gas.—This gas is obtained by passing steam through coke or anthracite coal when they are intensely heated. It consists chiefly of carbon monoxid, hydrogen, and small amounts of methane, carbon dioxid, nitrogen, and oxygen. One great disadvantage of water gas is that it is odorless and, consequently, its escape from a gas pipe may escape detection. This gas, owing to its large content of carbon monoxid, is excessively poisonous; carbon monoxid causing death by asphyxiation because it unites so strongly with the hemoglobin of the blood that it is not easily displaced and, consequently, the hemoglobin cannot unite with oxygen.

Natural gas and water gas burn with a blue flame and are suitable for illuminating purposes only when mixed with other gases.

Acetylene (C_2H_2). This gas, as stated in Chapter II., is, for practical purposes, produced by the interaction of water and calcium carbide. It is usually prepared in private plants for country houses that are away from the main gas and electric lines. It is used especially as an illuminant, but also for cooking and heating. Unfortunately, it is exceedingly explosive and not only forms explosive mixtures with air, but, when compressed, is itself explosive. When burned in the presence of oxygen, acetylene yields a flame hot enough to melt iron; thus, it is sometimes used in a specially constructed apparatus, instead of solder, for fastening metal joints together.

Gasoline gas.—This is another gas prepared in private plants for country places. It is made by allowing gasoline to volatilize in the presence of air, the combination being regulated so that an explosive mixture will not be formed in the pipes.

Compressed gas.—Certain of the hydrocarbon gases are compressed and stored in cylinders that can be shipped to country houses and institutions and attached to their gas-supply pipes. Gas so treated is called *compressed gas*.

Electricity

Electricity is rapidly replacing gas as an illuminant and, as soon as the current can be produced more cheaply, it will probably do so also for heating and cooking, for it has many advantages. Two of these are: The wires for its conveyance are cheaper to install than the gas pipes, especially when houses are far apart; and it does not add to the impurities of the air in buildings, as do other forms of illuminants. The way in which heat and light are obtained from an electric current was described in Chapter VIII.

CHAPTER XIV

SOLUTIONS. ACIDS. BASES AND SALTS

Different Kinds of Solutions—Different Kinds of Acids.—Characteristics of Acids—Tests for Acids—Properties of Bases and Alkalies—Neutralization—Different Kinds of Salts—Alkaloids and Their Salts—Nature of Fats—Saponification—Nature of Soaps.

Solutions

Definitions.—When the word solution is used as a verb it stands for *the dissolving of a substance in a liquid;* used as a noun, it implies a *liquid with some substance dissolved in it.* The liquid which is used to dissolve a substance is called the *solvent*, and the substance that is dissolved is termed the *solute.* An unlimited amount of solute cannot be dissolved in a solvent, and when the latter contains as much of a solute as it can hold in solution at ordinary temperatures it is said to be *saturated.* By heating a solvent, it can be made to take up more solid solute and, if it is made to take up so much that, when it cools, the addition of a little extra will cause some of that in solution to be precipitated, the solution is said to be *supersaturated.* Such a solution is very unstable; after it has cooled, even shaking the flask in which it is will precipitate the amount of solid that is in excess of the quantity required for saturation. Heat lessens, instead of increasing, the amount of gas that can be held in solution, because heat increases the molecular motion and expansion of gases more quickly and to a greater extent than that of liquids.

Nature of matter used for solutions.—Matter in any one of its physical states (*i. e.*, solid, liquid, and gas) can be used as a solute. There are, of course, many substances that can be dissolved only at extremely high temperatures, and all substances are not soluble in the same kind of solvents. Fat, for example, is not soluble in water or in cold alcohol, but it is soluble in hot alcohol, alkaline solutions, ether, benzene, and similar substances.

Water and alcohol are the solvents most frequently used in making solutions. A solution for which water is used as the solvent is called an *aqueous solution*, and one for which alcohol is used as the solvent is termed an *alcoholic solution*.

Isotonic, hypertonic, and hypotonic solutions.—These terms are often used in physiology and medicine when comparing the relative percentage of salts of any kind contained in the blood and in other solutions. By an isotonic solution is meant one that contains the same percentage of salts as the blood; by hypertonic, one with a higher content of salts than the blood; by hypotonic, one with a smaller amount of salts than the blood.

Acids

Classification.—Acids are classified as organic and inorganic acids. The former are derived from organic (*i. e.*, either animal or vegetable) matter, and the inorganic, from inorganic substances. As all organic compounds contain carbon, this element is always present in organic acids. Some of the more common important organic acids are as follows:

Acetic	acid	$CH_3.CO_2H$	Malic	acid	$C_4H_6O_5$
Butyric	"	$H.C_4H_7O_2$	Oleic	"	$H.C_{18}H_{33}O_2$
Carbonic	"	H_2CO_3	Oxalic	"	$H_2C_2O_4$
Citric	"	$H_3.C_6H_5O_7$	Palmitic	"	$H.C_{16}H_{31}O_2$
Formic	"	$H.CHO_2$	Stearic	"	$H.C_{18}H_{35}O_2$
Lactic	"	$H.C_3H_5O_3$	Tartaric	"	$H_2.C_4H_4O_6$

Acetic acid is derived principally from alcohol as the result of fermentation brought about by the agency of a minute organism commonly called *mother of vinegar*, or from the distillation of wood in the absence of air. It is formed also in decaying fruit from the fermentation of the fruit sugar. The sour taste of vinegar is due to the presence of acetic acid—good vinegar containing about 6 per cent. of the acid. The glacial form of the acid is produced by lowering the temperature until the acid crystallizes. It has a strong affinity for water and is therefore used as a dehydrating agent (*i. e.* to remove water).

Carbonic acid is a weak unstable acid that breaks up into carbon dioxid and water whenever it is set free from its salts.

Citric acid occurs in large quantities in the citrus fruits (*i.e.*, lemons, grape fruit, oranges, and limes) and slightly in quince, gooseberries, strawberries, raspberries, currants, and cranberries.

Formic acid is found in plants such as stinging nettles, in red ants, and in a secretion in the sting of the bee. It can be made by the oxidation of wood alcohol.

Lactic acid is formed from lactose (the sugar in milk) when, as the result of bacterial action, it ferments. The souring of milk is due to formation of this acid.

Malic acid and malates occur in apples, pears, currants, blackberries, raspberries, quince, pineapples, cherries, and rhubarb.

Oleic acid will be referred to in Chapter XVIII.

Oxalic acid is present in the form of salts in many plants, such as the sorrels. It can be made by the action of nitric acid on carbohydrates and by heating wood shavings or sawdust with caustic potash or soda.

Palmitic acid, see Chapter XVIII.

Stearic acid, see Chapter XVIII.

Tartaric acid and its salts occur chiefly in grapes.

Other important organic acids are the amino acids and the acid amides. These, of which there are a large number, are formed both in plants and in animals. In the latter they result from the digestion and the metabolism of protein substances.

The inorganic acids.—The more common important inorganic acids are:

Hydrochloric or muriatic acid (HCl).
Hydrobromic acid (HBr).
Nitric acid or aqua fortis (HNO_3).
Phosphoric acid (H_3PO_5).
Sulphuric acid or oil of vitriol (H_2SO_4).

Acids and hydrogen.—The molecules of all acids have at least one atom of hydrogen that is detached in chemical reactions, and it is upon this atom that their acid characteristics depend; losing this, they lose their acid properties. Some acids contain more hydrogen than the detachable atom or atoms, and when writing the formulæ for such acids the replaceable hydrogen is sometimes separated by a period—thus, $H.C_2H_3O_2$. Hydrogen unites with certain elements to produce acids; for instance, hydrogen and chlorin unite to form hydrochloric acid, hydrogen and bromin unite to produce hydrobromic acid, hydrogen and fluorin combine to form hydrofluoric acid.

Anhydrids.—Acids are formed also by the union of certain gaseous oxids and water. For example, nitric acid, sulphuric and sulphurous acid, carbonic acid, phosphoric acid, are formed by the union of water and oxids of the elements nitrogen, sulphur, carbon, and phosphorus respectively. Such oxids are often spoken of as *anhydrids*. *Anhydrids* are also obtained by the abstraction of water from acids.

Strong and weak acids.—Some acids are much more active chemically than others; *i.e.*, they react more quickly and with a greater number of substances to form new compounds. Those which do are classed as *strong acids* and those which do not as *weak acids*. The important strong acids are hydrochloric, nitric, sulphuric, and oxalic. Tartaric and citric are moderately strong; and acetic, carbonic, boric, and the fatty acids—as oleic, stearic, etc. —are weak.

Properties of acids.—All acids, whether organic or inorganic, have certain marked characteristics, viz.: they have a sour taste; they are very soluble in water; they contain hydrogen in loose combination and part with it when they come in contact with substances for which they have an affinity. On account of this last characteristic, acids decompose and often dissolve many substances, and they form different compounds in the process. They act in this way upon many metals, as will be seen in the paragraphs devoted to salts and in Chapter XVI. In fact, acids have such a strong affinity for basic substances in nature and the animal body that they occur more frequently in the form of salts than in the free state. The strong acids are very corrosive. Their action upon metals and fabrics will be discussed in Chapter XVI. When acids are perfectly free from water or when they are dissolved in liquids which do not dissociate them into ions, they do not show acid characteristics.

Tests for acids.—Certain substances that have been named *indicators* are used for the detection of acid matter. The more common ones are: Litmus, which is derived from a small lichen and can be obtained in liquid or powder form or as paper dyed with litmus; and methyl orange and phenolphthalein which are coal-tar products. Acids turn blue litmus and yellow methyl orange to red, and they make pink phenolphthalein colorless.

Nomenclature of acids.—Acids which consist of only one other element in addition to hydrogen are called *binary acids;* those which contain three elements are classed as *ternary acids.* Binary acids are given names consisting of the prefix *hydro*, the name of the second element, and the suffix *ic*—*e.g.*, hydro-chlor-ic acid. Oxygen is the third element in the majority of acids, and many acids differ from each other in composition only in the amount of oxygen which they contain. When there are several acids similar to each other save in this respect, the best known one is given the suffix *ic*, *e.g.*, chloric, sulphuric, nitric, etc. If there is an acid of the same name containing less oxygen than that with the suffix *ic*, it has the suffix *ous*, instead of *ic*; if there is a similar acid with still less oxygen, it retains the suffix *ous*, but it also has the prefix *hypo*. If an element forms an acid with a greater amount of oxygen than the *ic* acid, the suffix *ic* is retained and it has also the prefix *per*. The following table of the chlorin acids illustrates this nomenclature:

Hydrochloric acid—HCl
Hypochlorous " $HClO$
Chlorous " $HClO_2$
Chloric " $HClO_3$
Perchloric " $HClO_4$

Bases and Alkalies

Definitions.—The terms bases, hydroxids, and alkalies are often used interchangeably, but many chemists do not consider this correct because, though bases are alkalies, there are some substances that have an alkaline reaction that are not true bases. A base or hydroxid is often defined *as a substance which produces hydroxyl ions when dissolved in water or other dissociating liquid;* and an alkali, *as any one of a class of compounds which forms salts with acids and soaps with fats.*

Reaction of the bases, etc.—The bases and alkalies have an alkaline reaction; *i.e.*, they change red litmus to blue, red methyl orange to yellow, and colorless phenolphthalein to pink.

Names of some of the important bases.—These are:

Ammonium hydroxid or ammonia water (NH_4OH).

Calcium hydroxid [$Ca(OH)_2$]. When in solution, this is known as lime water.

Potassium hydroxid or caustic potash or lye (KOH).

Sodium hydroxid or caustic soda ($NaOH$).

Derivation of the term hydroxid.—The hydroxids received their name because the elements which form them do so by their interaction with water. For example, if the element sodium is dropped on water, some of the latter will be decomposed and the metal will at once unite with the oxygen and half of the hydrogen, and the other part of the hydrogen will escape in gaseous form. The reaction takes place so quickly that sufficient heat is caused to melt the metal, but, if the remaining water is evaporated, the sodium hydroxid formed by the interaction will remain as a white solid. The reaction that occurs is as follows: $Na + H_2O = NaOH + H$.

Characteristics of hydroxids.—All true hydroxids contain the hydroxyl radical (OH) and the majority of them consist of a metal in combination with this radical. When bases are dissolved in water, they dissociate into two kinds of ions one of which is always the hydroxyl ion and the other the metal element, as described in Chapter VII. Bases have a soapy feeling, a brackish taste, and an alkaline reaction. They destroy tissues by abstracting their water, dissolving their albumin, and saponifying the fat. They destroy bacteria by the same means.

The action of alkalies upon albumin and fat makes them valuable for cleansing agents, but, as will be seen in Chapter XVI., they, like acids, cannot be used for cleaning all kinds of substances, and when strong basic solutions are used on absorbent material, their action must be neutralized by the use of an acid.

Neutralization

Experiment 19. Object: To study neutralization.

Articles required: An iron stand or tripod, a

Bunsen burner, wire netting, an evaporating dish half full of water, NaOH, HCl, litmus paper.

Dissolve a small piece of NaOH in an evaporating dish half full of water. Add dilute HCl drop by drop until the solution will slightly redden blue litmus paper. Place the dish on a wire netting, over the flame and allow its contents to evaporate to dryness. Continue the heating until the yellow color disappears—this being due to the excess HCl that was added. Add a little warm water and evaporate this in order to remove all traces of the acid. Test the residue with litmus. Is it acid, alkaline, or neutral (*i. e.*, neither acid nor alkaline)? Taste the residue; what is it?

Write the equation showing the reaction.

The two chemicals used for this experiment were corrosive poisons, but, when they united, the acid lost its hydrogen atom and the base its hydroxyl radical, these uniting to form water, while the Na and the Cl united to form a salt that was neither acid nor alkaline and, therefore, neutral. Thus: $NaOH+HCl=NaCl+H_2O$. This operation is called *neutralization*.

As alkalies and acids neutralize each other, acids are used in the treatment of poisoning by alkalies and to remove stains made by alkalies, and alkalies are used to counteract the action of acids.

Salts

Salts and salt-like substances compose a very large proportion of the compounds forming the earth's crust; they enter into, and are essential for, animal and plant tissues; in fact, they are found in great varieties everywhere in nature.

How salts are formed.—Salts may be formed by the addition of an acid to (1) a metal, (2) the oxid of a metal, (3) an hydroxid, (4) an alkaloid. For example:

Metal		*Acid*		*Salt*		*By-Product*
Sodium		Hydrochloric		Sodium chlorid		Hydrogen
Na	+	HCl	=	NaCl	+	H
Zinc		Sulphuric		Zinc Sulphate		
Zn	+	H_2SO_4	=	$ZnSO_4$	+	H
Magnesium		Sulphuric		Magnesium sulphate		
Mg	+	H_2SO_4	=	$MgSO_4$	+	H
Metal oxid		*Acid*		*Salt*		*By-Product*
Ferrous oxid		Hydrochloric		Ferrous chlorid		Water
FeO	+	2HCl	=	$FeCl_2$	+	H_2O
Hydroxid		*Acid*		*Salt*		*By-Product*
Sodium hydroxid		Sulphuric		Sodium sulphate		
2NaOH	+	H_2SO_4	=	$2NaSO_4$	+	$2H_2O$

Alkaloids and their salts will be discussed later.

How salts can be dissociated.—Salts are not only being constantly formed in nature, even in the animal body, but they are also being dissociated, and they often form new compounds after dissociation. How easily decomposition occurs when a salt comes in contact with a substance that has a stronger affinity for one of its elements than those with which it is combined, can be seen by the following experiment:

Experiment 20. Articles required: A flask; two corks, one with one hole and the other with two holes through the center; a thistle tube; a piece of glass tubing bent as in Fig. 58; a bottle one third full of water; sodium chlorid (NaCl); sulphuric acid (H_2SO_4); litmus paper.

Procedure: Put about 16 grams of NaCl into the flask, connect the apparatus as in Fig. 57. Through the thistle tube, pour some H_2SO_4 upon the salt. After action ceases, test the water with litmus paper.

The water, it will be found, is acid, because the chlorin was liberated from the sodium and the hydrogen from the sulphuric acid, and these two gases, which passed through the tube to the water, united, forming HCl. The liquid in the flask contains a new salt, and it can be obtained by evaporating the liquid. The salt is sodium sulphate (Na_2SO_4).

Write the equation showing what happened in the reaction.

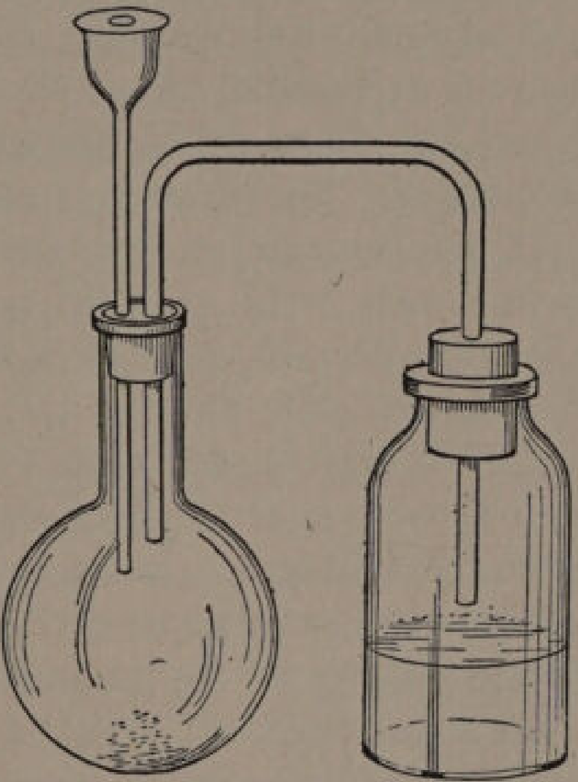

FIG. 57. APPARATUS USED FOR EXPERIMENT 20.

Salts are also dissociated to some extent when dissolved in water, as shown in Chapter VII. and the decomposition is hastened when an electric current is passed through the solution. The dissociation of many salts is also hastened when their solutions are heated. It is necessary for nurses to remember this. An example of a very common error arising from ignorance of the fact is the sterilizing of sodium bicarbonate solutions. When this is done, the patient gets very little sodium bicarbonate, the salt having been decomposed into its ions during the sterilization. If, however, as is done in many hospitals, the quantity of sodium bicarbonate required for each flask of solution is wrapped in a separate package, sterilized by *dry heat* at about 60° C. on three successive days, and added to the sterile, distilled water shortly before use, the solution will be sterile and the soda will not be dissociated to any extent.

Normal, acid, and basic salts.—Salts which have neither acid nor alkaline characteristics are called neutral or normal salts. There are, however, some salts that have acid and some that have alkaline properties. For example, any acid that has more than one replaceable hydrogen atom may be made to form a salt with an acid reaction; thus, $H_2SO_4 + 2NaOH = Na_2SO_4 + 2H_2O$. Sodium sulphate ($Na_2SO_4$) is a normal salt; but the salt formed thus, $H_2SO_4 + NaOH = NaHSO_4 + H_2O$, is an acid salt, for, as can be seen by looking at the formula, $NaHSO_4$ (acid sodium sulphate) still retains an atom of hydrogen, since there were not sufficient hydroxyl ions provided to unite with all the hydrogen. In the same way, if a base, such as magnesium hydroxid [$Mg(OH)_2$], and an acid, as HCl, are brought together in such quantity that there will be an equal number of molecules of each, there will be only half enough hydrogen for the hydroxyl ions present, there being, as can be seen by the formula, two hydroxyl ions in each molecule of the base, and the result will be as follows: $Mg(OH)_2 + HCl = MgOHCl + H_2O$. MgOHCl is called *basic magnesium chlorid.* When the HCl is in the proportion of $Mg(OH)_2 + 2HCl$, the reaction is $MgCl + 2H_2O$, and MgCl—magnesium chlorid—is a normal or neutral salt.

Naming of salts.—(Read the names of the acids, page 197.) The names of the salts are derived from those of the bases or metals, etc., and of the acids which produced the salts. As in the case of the acids, the suffix usually shows the lack of oxygen or the amount of oxygen present. A salt derived from a binary acid generally has the suffix *id* or *ide*, thus chlorids are formed from hydrochloric acid, bromids from bromic acid, etc. The salts of a ternary acid ending in *ous* have the *ous* changed to *ite*, the salts from a ternary acid with a name ending in *ic* have the *ic* changed to *ate*; thus: nitrous acid produces nitrites, nitric acid, nitrates; sulphurous acid, sulphites, sulphuric acid, sulphates; acetic acid, acetates.

Class exercise.—Write the equations showing how the following salts are formed:

Sodium chlorid.
Potassium bromid.
Acid sodium carbonate (bicarbonate of soda).
Carbonate of soda.
Magnesium sulphate.
Calcium nitrate.
Silver nitrate.
Copper sulphate.
Silver chlorid.

Alkaloids and Their Salts

The term alkaloid is derived from two Greek words signifying *alkali-like*, and the name was first given to certain unsaturated ammonia compounds of vegetable origin which, like the alkalies, combine with acids to form salts.

It was once thought that alkaloids were formed only in plants, and when they were found in a post-mortem examination it was decided that the individual had been poisoned by one of the poisonous alkaloids derived from certain plants. It is now known, however, that many bacteria produce alkaloids—ptomaines—in animal matter and that, after decomposition of the body has set in, alkaloids from this source may be found in animal organs. Some of the important vegetable alkaloids are morphin, strychnin, quinin, caffein, nicotin. The majority of alkaloids are insoluble in water, which is one reason why their salts, and not the free alkaloids, are used as medicine. Morphin sulphate is made from the alkaloid morphin and sulphuric acid; strychnin sulphate, from the alkaloid strychnin and sulphuric acid; iron and strychnin citrate, from iron, strychnin, and citric acid.

Esters or Ethereal Salts

How esters are formed.—Whenever an acid acts upon an alcohol, the acid is neutralized, wholly or in part, and a substance analogous to a salt is formed. Thus nitric acid and ethyl alcohol combine to form ethyl nitrate:

$$\overset{\text{Alcohol}}{C_2H_5OH} + \overset{\text{Nitric acid}}{HNO_3} = \overset{\text{Ethyl Nitrate}}{C_2H_5NO_3} + \overset{\text{Water}}{H_2O}$$

and acetic acid and ethyl alcohol when combined form ethyl acetate, thus:

$$C_2H_5OH+CH_3COOH=CH_3COOHC_2H_5+H_2O$$

Fats

Nature of simple fats.—Fats, which in chemistry are often spoken of as glycerids, are esters of glycerin or, as it is often called, glycerol, and certain organic acids, known as fatty acids. The most common of these acids are stearic, palmitic, oleic, and butyric. Stearic acid is solid, palmitic semi-solid, and oleic and butyric are liquid at ordinary temperatures. The fats which they form are known as *stearin*, *palmitin*, *olein*, and *butyrin*. Glycerol, as shown on page 166, being a triatomic alcohol, has three OH radicals that can be replaced, and as fatty acids are monatomic, having only one replaceable hydrogen atom, a molecule of any one of the glycerids will consist of one molecule of glycerin and three of fatty acids. For example:

$$\overset{\text{Glycerol}}{C_3H_5(OH)_3} + \overset{\text{Stearic Acid}}{3C_{18}H_{36}O_2} = \overset{\text{Stearin}}{C_3H_5(C_{18}H_{35}O_2)_3} + \overset{\text{Water}}{3H_2O}$$

Fats will be further discussed in the chapters devoted to food.

Saponification

When fat is boiled with an alkali, the glycerin separates from the fatty acids and these unite with the alkali, thereby forming soap. The process is called *saponification.*

Saponification occurs also in the animal body, in the small intestine under the influence of an enzyme produced in the pancreas and the alkaline salts of the bile and the intestinal and pancreatic juices.

Soaps, being formed by the interaction of acids and a base, are analogous to salts. They will be further considered in Chapter XVI.

CHAPTER XV

WATER

City and Country Water Supplies—Ground Water—Springs—Wells—Classification of Foreign Substances Found in Water—Methods of Purifying Water—Hard and Soft Water—Methods of Softening Different Kinds of Hard Water—Objections to the Use of Hard Water.

Water Supply

The city water supply.—A city's water supply is brought by aqueducts or pipes from the nearest suitable river or lake.

The water is forced into the pipes in different ways. If the city is near a hill on which there is a lake of sufficient size, at a high enough level to be above the houses in the city, the pipes are arranged so that the water will flow from the lake to the city by gravity, and the pressure upon it by that in the lake is sufficient to force it from the main pipes into the house pipes and, if desired, up to a height almost to the level of the water in the lake.

When there is no available lake, a common method is to build a reservoir on some high land near the city and force the water, by means of pumps, from its lower plane to the reservoir, from which it flows to the city in the same way as from the lake. It is often necessary to build an aqueduct or to lay pipes between the site of the reservoir and the source of the water supply.

In flat countries, the water is often pumped directly into the main pipes at the source of the supply.

Houses built on heights are often provided with supply tanks or, as they are often called, *standpipes.* This is because it requires a great deal of pressure to raise a large supply of water to a high level, and if, when there is not much water being used, any excess supply passes into a reservoir from which it will flow when the amount being provided is insufficient, it minimizes the amount of pressure that will be necessary to provide an adequate supply at all times.

The country water supply.—If people living in the country, away from the line of the main pipes, want running water in their houses, it must be pumped directly into the pipes supplying each house. One form of pump used for this purpose is the windmill pump, in which the mechanism that produces the pressure and suction necessary to raise the water is operated by the wheel at the top of the mill, and the wheel is driven by the wind. As a rule, the water is raised to a tank that is placed considerably higher than the faucets through which it is to flow and with which it is connected by pipes, and thus gravity and the pressure of the water in the tank are the two chief forces driving the water through the pipes. Pumps that are operated by motors driven by electrical power or power derived from the combustion of gasoline or other fuel are now much used instead of windmill pumps, when the expense of the fuel need not be considered. The water thus pumped into country houses may be obtained from a neighboring lake or river, but where such does not exist a well is the usual source of supply.

Ground water and wells.—In order to understand how wells and springs can exist, even in countries where rain does not fall for months at a time, it is necessary to know that the crust of the earth is composed of several different kinds of material, some of which are very porous and others perfectly impervious to water. In many places, these different substances have become deposited in uneven layers or strata with a varying depth of porous material at the top. The water of rain and melting snow sinks through the porous strata until it comes to a stratum of rock or thick hard clay that prevents its passage and then it flows or filters along the surface of this bed until it either finds an outlet into a stream, lake, or other body of water or else it forces its way to the surface, thus giving rise to a spring; the common cause of this is obstruction to the flow of the water by a mass of impervious material. The water flowing in this way underground is called *ground water* and its surface is known as the *ground-water level.* The depth of the ground water will depend upon the amount of rain or snow that falls.

In many parts of the world a non-porous stratum dips and then rises, forming underground valleys or basins, and the ground water flows into and fills them. These are known as *artesian basins* and a well cut into an artesian basin is known as an *artesian well.* If the non-porous material forming the basin is higher at one side than the other and extends near enough to the surface of the ground, the water will force an exit, thus giving rise to an *artesian spring.* The size of artesian basins varies greatly; some are quite small, but some are known to be at least a hundred miles in circumference and very deep.

Wells can be made where there is no artesian basin, if there is a plentiful supply of ground water at a moderate depth. It is usual in making such wells to dig somewhat below the surface of the non-porous stratum.

Driven wells are now more used than the old-fashioned *open well*. A driven well is made by inserting a pipe, that has perforations in its lower end, into the ground to such a depth that its point is embedded in the non-porous stratum and its perforations are surrounded by ground water. The upper end of the pipe is connected to a pump. This may be a simple hand pump, or a mill-wheel pump, or a motor-driven pump.

Composition of water.—Pure water, as its chemical symbol—H_2O—shows, is an oxid of hydrogen. Water is, however, seldom found pure in nature for, being an excellent solvent, it dissolves and absorbs substances from the soil and rocks through and over which it flows, and even rain water holds in solution substances absorbed during its passage through the air.

Foreign substances likely to be found in water.—These may be both organic and inorganic and they are usually considered under three headings: (1) Those likely to be injurious to the health; (2) those which render water unfit for cleansing and culinary purposes; (3) those used for medicinal purposes.

It is the organic matter that has usually to be considered with relation to the prevention of disease; for when there is a large amount of organic matter present, there is usually a large number of bacteria, and some of these may be pathogenic. As a rule, the inorganic matter in water has no harmful effect upon the health; in fact, its presence in drinking water is usually considered an advantage, but, as will be seen later, its presence in more than small amounts in water used for cooking and cleaning, especially the latter, is a very great disadvantage.

Bacteria in water.—The water of streams, rivers, lakes, ponds, and even that of oceans contains varying numbers of bacteria. Rain water may contain a very large number that it has gathered with the dust in its passage through the air. The water of wells varies very greatly in the number of bacteria it contains; that from a deep driven well may be practically bacteria free, since the ground through which the water has passed has acted as a filter, but shallow wells, especially those situated near stables, privies, or where the drainage from sinks has access, are likely to be terribly contaminated. For the bacteria will be carried down and for some distance through the soil with the rain water and other fluid matter.

Kinds of bacteria in water.—The number of different species of bacteria occurring in water is very large. Some varieties seem to live on the surface of water as their natural habitat, but, as far as is known, these are not pathogenic and neither are the several varieties of soil bacteria found in streams and near the borders of lakes and rivers, especially after rain. The pathogenic bacteria most frequently found in water are those which are habitants of the intestine, and they occur most frequently in water that receives sewage. Many epidemics of typhoid and cholera have been traced to such a source.

Methods of purifying water.—Large supplies of water are usually purified either by storing the water in reservoirs or by filtration.

Sedimentation.—When water is stored in reservoirs, the solid matter, including bacteria, tends to fall to the bottom, and even pathogenic bacteria soon die for there is nothing in the water to provide them with food. Also, the sun's rays, to which the water is exposed, acts as a disinfectant. This process of self-purification by sedimentation is constantly going on also in rivers and lakes, and it has been found that rivers into which the sewage of a city is emptied may be free from pathogenic bacteria several miles from the city though highly contaminated within a considerable radius of it.

Filtration.—By filtration is meant the act or process of passing liquids through porous solids. This process is very frequently used for the purification of both large and small supplies of water. It is often used in the form of filter beds in connection with reservoirs.

A variety of filter bed frequently used consists of successive layers of stones and of coarse and fine gravel and sand; the whole bed being 1 to 2 meters thick. Such a bed is often connected by pipes with a reservoir.

Another method of filtering frequently used is the passing of water through filters of unglazed porcelain, such as the Pasteur-Chamberland and the Berkfeld filters. These two varieties of filters are very effective, if they are frequently baked and kept clean, otherwise they are useless. Before purchasing unknown kinds of filters for household use it is well to have water which has been passed through them tested, for many filters offered for sale are quite worthless.

Sterilization.—Though a good filter will remove nearly all germs present in water, filtration is not to be depended upon to sterilize water that is to be used for surgical work or, when the presence of pathogenic bacteria is suspected, for drinking. To be sterilized, water must be boiled.

Distillation.—Filtration and sterilization will rid water of bacteria, but neither process will remove the salts that are in solution in the water. If the water is distilled, however, nothing with a higher boiling point than water will distill over, and thus the salts will be left behind. For certain purposes, *e. g.*, dissolving drugs and chemicals, it is often necessary to use distilled water, since some of the salts present in undistilled water, may unite with a drug or chemical and entirely change its nature. To prove this, add about ten drops of 2% silver-nitrate solution to some distilled water; do likewise to some tap water. The distilled water will remain clear, but the tap water usually becomes more or less cloudy, because water that has not been distilled generally contains some chlorides, and the following reaction occurs between the chlorid and nitrate salts:

$$\overset{\text{Silver nitrate}}{AgNO_3} + \overset{\text{Sodium chlorid}}{NaCl} = \overset{\text{Silver chlorid}}{AgCl} + \overset{\text{Sodium nitrate}}{NaNO_3}$$

The silver chlorid thus formed is insoluble in water, hence the clouding of the latter. It is for this reason that distilled water must be used for diluting the silver preparations often used for irrigations in the treatment of gonococci infections, silver chlorid being absolutely useless for the purpose.

Experiment 21. Object: To demonstrate the effect of distillation as compared with filtration.

Articles required: 2 stands; 2 flasks, one half full of water; a cork with a hole in the center; glass tubing bent as in Fig. 58; a dish with ice in it; wire gauze; a

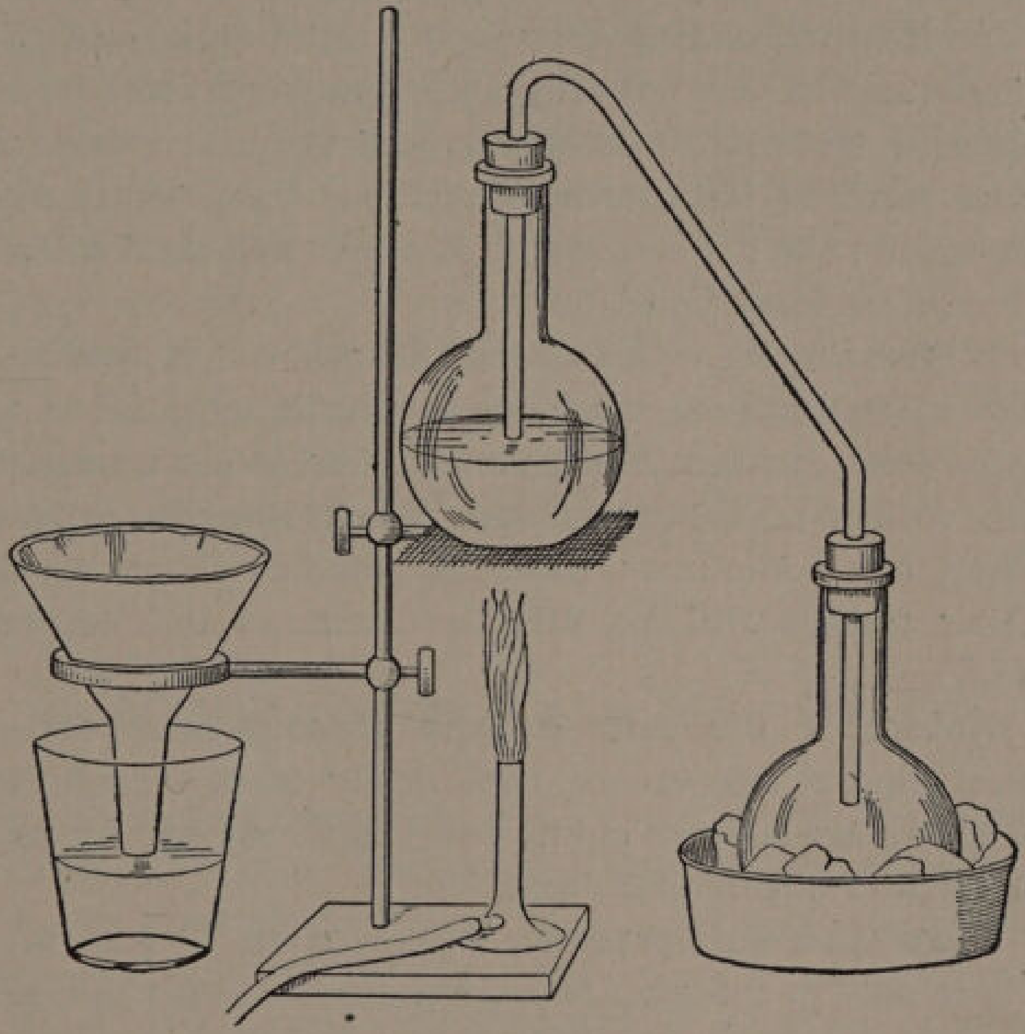

FIG. 58. APPARATUS ARRANGED FOR DISTILLATION AND FILTRATION.

Bunsen burner; a funnel; filter paper; 2 beakers, one three quarters full of water; copper sulphate.

Procedure: Color the water in the flask and that in the beaker an equally bright blue by the addition of some copper sulphate. Line the funnel with filter paper and suspend it above the empty beaker. Pour the copper sulphate solution into the funnel, slowly, so as not to tear the paper, and allow it to filter.

Arrange the distilling apparatus as shown in Fig. 58. Allow the water to boil. After it has done so for some time, it will be noticed that the vapor is passing through the tube and condensing in the empty flask. (The ice surrounding this flask is used to hasten the condensation by lowering the temperature.) The distilled water, it will be noticed, is quite colorless, since nothing but the steam (H_2O) passed into the flask, but the filtered water is about as bright a blue after, as before, filtration.

Mineral water.—Water that flows through localities rich in mineral matter, such as lithia, sulphur, iron, will often contain enough of the minerals to flavor the water; such waters are called *mineral waters.* They are sometimes used for medicinal purposes. Their nature will be further discussed in Chapter XIX.

Nature of the salts usually present in water.—The mineral substances that occur most frequently in water are sodium chlorid and different compounds of calcium and magnesium.

Hard and soft waters.—When soap is put into water part of the soap is decomposed, and if there are lime or magnesium salts in the water they will unite with the fatty acids set free by the decomposition and form a lime soap which, being insoluble in water, is precipitated in the form of a hard curd-like substance. The soap will form no lather and, consequently, will be of no use until enough has been used to combine with all the calcium or magnesium present. Waters which contain salts of lime or magnesium are, therefore, said to *destroy soap* and they are called *hard waters;* waters which do not contain these salts or only very small amounts of them are said to be *soft.*

How hard water can be softened.—Sometimes the lime and magnesium in water are in the form of bicarbonates (page 202) and sometimes in the form of sulphates, and as both the bicarbonates and sulphates of lime and magnesium are soluble in water they remain in solution. Consequently, in order to remove them and thereby soften the water, it is necessary to do something that will render the salts insoluble so that they will be precipitated.

Water that contains bicarbonates can be softened by boiling because during the boiling the salts lose CO_2 and are changed from soluble bicarbonate salts to insoluble carbonates; thus:

$$\underset{\text{bicarbonate}}{\overset{\text{Calcium}}{Ca(HCO_3)_2}} \quad + \text{boiling} \quad = \quad \underset{\text{carbonate}}{\overset{\text{Calcium}}{CaCO_3}}+CO_2+H_2O$$

The calcium carbonate thus formed, being insoluble in water, is precipitated and the precipitate is deposited upon the floor and sides of the utensil in which the water is boiled. Water that can be softened by boiling (*i. e.*, water the hardness of which is due to bicarbonates of lime or magnesium) is called *temporary hard water*.

Hardness due to sulphates of lime, etc., cannot be overcome by boiling, and the hardness therefore is spoken of as *permanent*, though, in reality, it is not permanent, for the salts can be easily changed to insoluble ones, and thus precipitated, by the use of sodium carbonate, borax, ammonia, and similar substances. When sodium carbonate is used the following reaction takes place:

Calcium sulphate		Sodium carbonate		Calcium carbonate		Sodium sulphate
$CaSO_4$	+	Na_2CO_3	=	$CaCO_3$	+	Na_2SO_4

The calcium carbonate thus formed, being insoluble, is precipitated; the sodium sulphate, being soluble, remains in the water, but it has not the objectionable qualities of the calcium sulphate.

Temporary hardness can be rectified in the same way as permanent, as well as by boiling; thus:

$$Ca(HCO_3)_2 + Na_2CO_3 = CaCO_3 + 2NaHCO_3$$

The softening of water for use in laundries.—Carbonate of soda is the agent usually used to soften water in laundries for the general washing, and if the soda is dissolved before being allowed to come in contact with the clothes and no more used than required to combine with the salts present in the water, it will do some fabrics no harm, but, if the soda is added to the clothes in solid form or in such amount that some of it remains uncombined in the water, it will adhere to the clothing and, as will be seen in Chapter XVII., this weakens and soon destroys the fabric.

Ammonia, though too expensive for general use, is often employed to soften water for the washing of fine fabrics, for, as it is volatile at ordinary temperatures, even if it is used in excess, it does not remain in the meshes of the fabric and, therefore, it is much less likely to injure material than soda. Borax, also, injures fabric less than soda.

Objections to the use of soap for softening water.—Water can of course be softened without the addition of an alkali, if a large amount of soap is used, but there are two great objections to this: (1) the insoluble calcium soap that is formed adheres to the material being washed and discolors it; (2) a much larger quantity of soap must be used than of soda and the soap is more expensive than soda. It has been estimated that it requires from 2 to 2½ ounces of soap for every gallon of water that has one degree of hardness[1] and that about .75 grams of soda will be the equivalent of this amount of soap. As water frequently has from 5 to 10 degrees of hardness and soda is usually only about one-third the price of laundry soap, it will be seen that if large quantities of water are used the difference in cost will be considerable.

Softening of municipal water supplies.—Municipal water supplies in districts where the water is unusually hard, are often softened by adding lime to the water. How it does so can be seen in the following equation:

$$\overset{\substack{\text{Calcium}\\\text{bicarbonate}}}{Ca(HCO_3)_2} + \overset{\substack{\text{Calcium}\\\text{hydroxid}}}{Ca(OH)_2} = \overset{\substack{\text{Calcium}\\\text{carbonate}}}{CaCO_3} + \overset{\text{Water}}{2H_2O}$$

Objections to the use of hard water.—Some important objections to the use of hard water other than the waste of soap and harm to clothing already referred to are: (1) The calcium soap, formed by the union of the fatty acids of the soap and the lime of the water, adheres to the sides of the sinks, utensils, etc., and requires time and frequently the use of some substance in which calcium soap is soluble, as kerosene, to remove it; also, it helps to block the sewer pipes into which the water is emptied. (2) In cooking

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[1] A water is said to have one degree of hardness when it contains, per gallon, as much calcium as there is in one grain of calcium carbonate. It is said to have two degrees of hardness when it has twice this amount, and so on.

with hard water, a deposit of lime salts forms upon the food and often hardens it and, in the case of tea and broths, this prevents the withdrawal of the extractives. It has been found in some instances that, using the same amount of tea leaves, ten ounces of tea made with fairly soft water is as strong as eighteen ounces brewed with hard water. Also, the lime unites with the legumin in peas and beans and makes it so hard and insoluble that the vegetables cannot be softened. (3) The fur-like coating which forms on boilers and kettles in which hard water is used is a very poor heat conductor; consequently, it takes longer to heat water in utensils so coated and this necessitates waste of fuel; besides, there is the expense of having the boilers cleaned to be considered and the damage to the boilers, etc., by the overheating of the metal which occurs when the heat is not carried away from it by the water; then there is the danger of an explosion should the coating crack suddenly, for, in such case, there will be a sudden, forcible burst of a

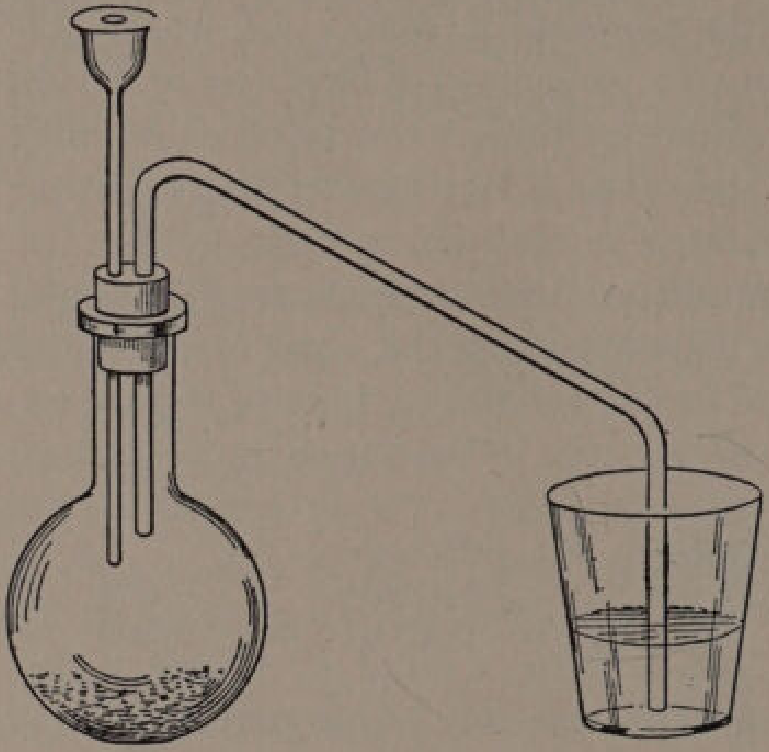

FIG. 59. APPARATUS ARRANGED FOR EXPERIMENT 22.

large quantity of steam which will be generated when the water comes in contact with the highly heated metal.

Experiment 22. Object: To show differences in hard waters.

Articles required: Apparatus for generating CO_2—viz., a flask, a thistle tube, a cork with two holes in it, a piece of glass tubing bent as in Fig. 60;—nitric acid; calcium carbonate; calcium sulphate; lime water; distilled water; 2 beakers.

Procedure: (*a*) Arrange the apparatus for generating CO_2 as in Fig. 59, put about 30 grams of calcium carbonate into the flask, place the free end of the tube in a beaker half full of lime water; pour some nitric acid, slowly, through the thistle tube. The following reaction then occurs:

$$\overset{\substack{\text{Calcium}\\\text{carbonate}}}{CaCO_3} + \overset{\substack{\text{Nitric}\\\text{acid}}}{2HNO_3} = CO_2+H_2O + \overset{\substack{\text{Calcium}\\\text{nitrate}}}{Ca(NO_3)_2}$$

The CO_2 passes into the lime water $[Ca\ (OH)_2]$. What is the white precipitate? Write the equation. Let the CO_2 continue to flow into the lime water until the latter becomes perfectly clear. Why does it become clear?

Write the equation.

(*b*) Prepare permanent hard water by dissolving 0.1 gram of calcium sulphate (plaster) in 500 c. c. of distilled water; generate CO_2 as in (*a*). Why are the results not the same?

Experiment 23. Object: To show some methods of softening water and of estimating the amount of soda necessary to use for the purpose.

Articles required: Soap solution made by shaving

50 grams of white Castile soap, dissolving it in 1 liter of hot water and filtering it; 2 graduated c. c. measures; 2 small flasks or bottles; calcium sulphate; distilled water; tap water that has been boiled for 15 minutes or more, and some that has not been boiled.

(*a*) Pour about 100 c. c. of the unboiled tap water into a small flask, add some soap solution, 1 c. c. at a time, and, between each addition of soap, shake the bottle and let it stand for one minute. Do this until a lather that will persist for one minute is formed. Note the quantity of soap that is used before this occurs.

Repeat the experiment with some of the boiled water and notice the difference in the quantity of soap used.

Experiment 24. Object: To estimate the amount of soda that it will be necessary to use to soften water.

Procedure: Add: (1) a few drops of saturated solution of sodium carbonate (20%) to 100 c. c. of the unboiled water, or, if the water is not hard, water in which calcium sulphate has been dissolved, as in Experiment 23; and (2) 1 c. c. of soap solution; shake the flask and let it stand for one minute as in Experiment 23. If a lather is not formed, take more water, and add a slightly larger amount of the soda solution, and repeat the procedures. Fresh water must be taken for each trial as the soap that is used would assist in the softening.

If the quantity of soda necessary to use to soften 100 c. c. of water is known it will be an easy matter to reckon the amount needed for any number of gallons.

Anyone having the supervision of a laundry should know just how much soda per gallon of the water used in their laundry is required to produce a soapsuds, without waste of soap, and how many gallons of water is used in their washers. Knowing this, it is an easy matter to estimate the quantity of soda required and to regulate the amount used and thus prevent unnecessary destruction of clothing.

Examples of the Different Ways in which Water is Held in Combination

Water of crystallization.—Certain solids, when dissolved in water, will, if the resulting solution is allowed to evaporate, separate out in the form of crystals. Such crystals, when heated, will give up a definite amount of water and in so doing lose their crystalline character; thus showing that their crystalline appearance was due to water. The water which enters into combination with solids in this way is spoken of as the *water of crystallization*.

Crystals which do not give up water when heated —quartz, for instance—do not owe their crystalline character to water.

Mechanically enclosed water.—Some crystals hold water mechanically enclosed and not in direct combination. To see the difference of the manner in which the two kinds of crystals part with their water, perform the following experiment:

Experiment 25. Procedure: Place some crystals of copper sulphate in an evaporating dish and some sodium chlorid in another dish and heat both of these over Bunsen burners. The crackling sound that occurs when the sodium chlorid becomes hot is due to the bursting of the crystals as the water held within them is changed to steam and expands. Such crystals are said to *decrepitate*. The water in the sodium crystals is combined mechanically, that in the copper sulphate is water of crystallization.

Examples of formulæ representing water of crystallization.—The amount of water that certain crystals hold mechanically is not always the same, but the amount of water of crystallization is always definite for each substance that forms crystals in this way; *e. g.*, copper sulphate and water combine in the ratio of 1 molecule of the copper sulphate to 5 molecules of water; therefore the formula for its crystals is $CuSO_4.5H_2O$. Calcium sulphate combines with water to form gypsum in the ratio of 1 part of the salt to 2 of water, therefore the formula for gypsum is $CaSO_4.2H_2O$.

Efflorescence. Deliquescence

Some crystals give up their water of crystallization very readily, even exposure to the air resulting in such a loss. Such compounds are said to be *efflorescent*. Other substances have exactly the opposite quality; they will, if exposed to the air, absorb moisture from it. Such compounds are said to be *deliquescent*.

Experiment 26. Place some crystals of sodium sulphate in an evaporating dish; in another dish place calcium chlorid. Leave these exposed to the air for several hours and report the result.

Because of its deliquescent property, calcium chlorid, in moist climates, is sometimes put, in open dishes, into cupboards where there are instruments or other articles that are injured by moisture.

Hydrated. Dehydrated. Anhydrous.—Hydrate and hydrated are other terms applied to hydroxids and to substances that contain water of crystallization, and the terms *dehydrated* and *anhydrous* are often applied to any substance from which water has been abstracted; *e. g.*, anhydrous alcohol.

CHAPTER XVI

THE CHEMISTRY AND METHODS OF CLEANING

Source and Composition of Some Common Detergents—Nature and Origin of Material Used for Utensils and the Action of Cleansing Agents in Common Use on these and on Paint, Varnish, and Wax.

Classification of substances used for cleaning.—According to the nature of their action, the majority of substances used for cleaning may be classified as solvents, saponifiers, and bleaching agents. Cleaning is effected also by mechanical means; *e. g.*, friction or rubbing, absorption and suction.

Solvents

By a solvent is meant a liquid that can dissolve and absorb matter of any kind. All substances are not soluble in all kinds of liquids and there are some compounds that cannot be dissolved at all under ordinary conditions. To remove soiling matter that is not readily soluble, it is often necessary to add something to it that will change it to a soluble substance; *e. g.*, the addition of an alkali to fat in order to saponify it and thus make it soluble in water.

Effect of the temperature of solvents upon soiling

matter.—Matter is usually more soluble in hot than in cold liquids. Protein matter (see page 277), however, is an exception to this rule because heat coagulates protein; for example, it will be difficult to clean bottles and other utensils in which milk has stood if they are put into hot water before all residue of the milk has been removed by rinsing the utensils with cold water. Likewise, if woven fabrics, such as cotton, linen, etc., that are stained with matter containing protein—as blood—are put into hot water, the protein will become so hardened into the fiber of the material that it will be difficult to remove it. As there is always likely to be protein matter in the substance soiling clothing, bedding, and the like, such things are soaked in cold water before being subjected to hot water, and as there is usually fatty matter in such substance the soaking in cold water must be followed by washing in hot water—fat being insoluble in cold water—and the use of soap or an alkali will be necessary.

Source, nature, and use of some common solvents.—Water is the solvent in most common use. Alcohol is often used as a solvent for the removal of stains made by drugs, many of these being soluble in alcohol but insoluble in water. Turpentine, a volatile oil obtained from certain species of pine, will dissolve paint, varnish, and, to a slight extent, fat; it is therefore often used to remove stains made by such substances. Fat solvents frequently used are: ether, a very volatile, inflammable liquid obtained by the action of strong sulphuric acid upon ethyl alcohol; benzene, a colorless, volatile, inflammable liquid obtained by distillation from coal tar; benzin, a volatile, inflammable distillate of petroleum; gasoline, which is similar to, and derived from the same source as, benzin; carbon tetrachlorid, a non-inflammable coal tar product; organic acids, as citric acid, which is obtained from lemons, and oxalic acid, which is made chiefly from wood by heating sawdust and shavings with caustic potash or soda. Ammonia will dissolve the oxids of copper and nickel and is therefore often added to polishes used for cleaning these metals. Kerosene or coal oil, a distillate of petroleum, will dissolve several of the metal oxids, especially that of iron, therefore it is much used to remove rust from iron; it will also dissolve the lime soap formed, as shown in Chapter XV., when soap is put into hard water. Oxalic acid is another oxid solvent and both it and kerosene act also, to some extent, by reduction—*i. e.*, uniting with oxygen and thus taking it away from another substance.

Alkalies and Soaps

Reason for use.—Fat is the most common constituent of soiling matter and thus one of the most necessary procedures in general cleaning is to render the fat soluble in water so that it can be washed away. Alkalies do this by saponification and soap has the same, but less pronounced, effect as the alkalies, because, when it is dissolved in water, some of its alkaline constituent is set free. Soap and soap powders are more frequently used than the free alkalies for their action is sufficiently strong for ordinary purposes and they have not the harmful effect upon many substances that the stronger detergents have. Alkalies, however, are necessary for some purposes, as shown in the preceding chapter.

Nature and source of the alkalies most frequently used for cleaning.—The alkalies in most common use as detergents are: Ammonia, borax, sodium and potassium compounds.

Ammonia.—Ammonia water (NH_4OH) can be made by, under special conditions, running ammonia gas (NH_3) into water, but its common source is the ammoniacal liquor which is produced during the destructive distillation of coal in the making of gas.

Borax.—Borax is a compound of sodium, oxygen, and boron—a non-metallic element that resembles silicon. Borax deposits occur in various parts of the world, principally in the deserts. Borax is much used, not only for cleaning, but also in many industries. It is too expensive to use for general laundry work and scouring, but it is used for cleaning material that would be injured by the stronger alkalies.

Potassium and sodium carbonates.—The soft, light-weight, silver-white, metal elements, potassium and sodium, are never found free in nature, but their compounds are numerous and common. The compounds that are of special interest in connection with the study of detergents are the carbonates and hydroxids.

Potash or potassium carbonate has been in use since early in the middle ages, when it was discovered that after wood was burned, certain soluble salts were found in the ashes which, if boiled in pots with water until the latter evaporated, left a solid residue that had cleaning properties. This residue was called *pot-ashes*. Later, it became known as *potash*, *lye*, and *potassium carbonate*. Some of the potash used to-day is made in the old way, but its more common sources are the refuse left in making beet sugar, the grease that is taken from sheep's wool in the washing that precedes its being made into cloth, and natural deposits such as exist near Magdeburg, in Saxony. These deposits are thought to have been left after the evaporation of primeval seas.

Sodium carbonate was, in early times, prepared from seaweeds, but during the French Revolution the supply of seaweed was hard to get in places where the soda was usually prepared and as, at that time, wood was the only source of potash, it was difficult to procure a sufficient quantity of these detergents to supply the demand. The Academy of Paris, therefore, offered a prize of 25,000 livres for the discovery of a method of making sodium carbonate from sodium chlorid. In 1791, Le Blanc, a Frenchman, succeeded in doing this. He added sulphuric acid to sodium chlorid, thereby obtaining sodium sulphate, thus: $2NaCl+H_2SO_4=Na_2SO_4+HCl$. He then reduced the sulphate to a sulphid by heating it with carbon, thus: $Na_2SO_4+2C=2Na_2S+2CO_2$, and, finally, by heating the sulphid with calcium carbonate, he got the desired compound: $Na_2S+CaCO_3=Na_2CO_3+CaS$.

Sodium carbonate can now be obtained by easier and cheaper methods than the Le Blanc process, but nevertheless this is still extensively used in order to obtain the HCl, which is a by-product of the first reaction.

Potassium and sodium hydroxid can be made as described on page 198, but, for commercial purposes, they are prepared from potassium and sodium carbonate. Some of the several methods used in their preparation depend upon electrolysis. The oldest method, which is still extensively used, consists in treating the carbonate with calcium hydroxid, thus: $Na_2CO_3+Ca(OH)_2=CaCO_3+2NaOH$ or $K_2CO_3+CA(OH)_2=2KOH+CaCO_3$.

Soaps.—Soaps are made by causing the interaction of a fat with an alkali. When this is done, the fat separates into its component parts—*i.e.*, fatty acids and glycerin—and the acids unite with the alkali, thereby forming the soap. This interaction may be produced (1) by what is known as the *cold process*, in which the fat is combined with either sodium or potassium hydroxid and either churned or subjected to heavy pressure; (2) by boiling the fat with a strong solution of either sodium or potassium hydroxid. Soap made by this process is in solution in the water and glycerin, but it may be *salted* out by the addition of sodium chlorid to the solution, soap being insoluble in salt solutions. After the soap is removed, the glycerin can be freed from foreign substances. This is, in fact, the usual source of glycerin. Soap made by the cold process retains the glycerin in combination. Toilet soaps are often made in this way.

Reasons for differences in soaps.—The various differences in soaps depend chiefly upon (1) which alkali is used; (2) the nature of the fat used; (3) the addition of extra substances to the soap; (4) faulty manufacture of the soap.

Hard and soft soap.—Soaps made with sodium hydroxid are much harder than those made with potassium and are classed as *hard soaps*, while those made with potassium are called *soft soaps*.

Fats of soap.—All kinds of fat are used for soap and various oils, *e.g.*, fish oil, olive oil, cocoanut oil, and castor oil. Although, as stated in the preceding paragraph, soaps made with soda are harder than those containing potassium, the kind of fat used and the amount of water left in the soap cause considerable variation in the degree of hardness of even the soda soaps; *e. g.*, soaps made with olive and similar oils are softer and lather more easily than those for which tallow is used.

Substances frequently added to soap.—Various substances, some of which are harmless and others very objectionable, are often added to soap.

Some of the additions that, under ordinary circumstances, are harmless are perfumes; medicinal substances—as tar and salicylic acid; emollients—as glycerin, almond oil, and oatmeal; disinfectants—as carbolic; such detergents as borax, kerosene, and naphtha. These detergents increase the cleansing value of soap for laundry purposes and, if not present in excess, do not injure fabrics. Other detergents sometimes added to soap are sodium and potassium carbonate and rosin (the resinous substance that remains after the distillation of oil of turpentine from the fresh pitch of pine wood; it is known also as colophony). The presence of these substances in soap that is used for rough cleaning is not a disadvantage, but, as will be seen in the experiments in Chapter XVII., even dilute alkaline solutions will injure silk and wool and they will also destroy paint and varnish, and solutions that are at all concentrated will injure some of the metals. Rosin, if present in more than very small amounts, will leave yellow spots on fabrics. For these reasons and, also, because sodium carbonate and potash are cheaper than even cheap soaps, it is generally considered better to buy pure soaps and, when required, add a stronger detergent to the water, or, for scrubbing, to use a scouring powder.

Fillers.—By fillers are meant cheap substances with little or no detergent properties that are added to soap in order to increase its weight and hardness. The substances most commonly used for the purpose are the sulphates of calcium and potassium, silicates, and chalk. Naturally, fillers are objectionable, if for no other reason than that the buyer is paying for a substance that is valueless for the purpose for which it is bought; there is, however, another objection, viz., when such soaps are used for laundry purposes the fillers often leave an objectionable sand-like deposit in the clothes.

Substances that may be present in soaps as the result of faulty manufacture.—Definite proportions of fat and alkali must be used in soap-making or there will be uncombined in the soap some of whichever substance was present in excess. Also, the soap must be boiled the regulation time, if saponification is to be complete. Unless it is, there will be both free alkali and free fat in the soap. As already stated, a small amount of free alkali is not always objectionable, but free fat decidedly interferes with a soap's cleansing properties.

Water in soap.—Well made soaps do not contain more than about 25 per cent. water, but the cheaper grades, especially those containing fillers, often have a much larger amount. Such soaps are comparatively soft and dissolve very rapidly in water; therefore, to prevent waste, they should be dried before use. This is done by exposing the soap to the air until it loses its moist appearance.

Castile and olive oil soaps are made of either castor or olive oil and sodium hydroxid.

Floating soaps are made by beating the hot soap before it solidifies and thus making it lighter by incorporating air.

Transparent soaps of good quality are made by dissolving soap in alcohol, filtering the resulting solution, and allowing the alcohol to evaporate. The cheaper transparent soaps are generally made by a cold process and usually contain free alkali.

Glycerin soaps are nearly always made by a cold process but do not necessarily contain free alkali, like the cheaper transparent soaps.

Mottled soaps may owe their coloring to impurities, but that of the good grades is the result of the addition of coloring matter that does not affect the quality of the soap.

The so-called **green soap,** which is extensively used for cleaning the skin preparatory to disinfection, is made of potassium hydroxid and linseed oil.

The **tincture of green soap** is made by dissolving green soap in alcohol and adding oil of lavender flowers.

Experiment 27. Object: To test for the presence of adulterants in soap. Articles required: Several samples of soap; sulphuric acid; alcohol; acetic anhydrid.

For each test, use both a pure soap and at least one soap known to contain the adulterant for which the test is made. Compare all results with that obtained from the pure soap.

Test for sodium and potassium carbonate.—In separate evaporating dishes, place some small, thin slices of each of the soaps to be tested. To each, add a little sulphuric acid. Effervescence shows the presence of excess of carbonates.

Test for rosin.[1]—Dissolve the soaps in water; acidify the solution with sulphuric acid. Filter. Remove

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[1] Rosin is not considered an adulterant, if the soap is sold as a rosin soap.

the precipitate from the filter paper and dissolve it in acetic anhydrid. Add a few drops of this solution to about 2 c. c. of 50 per cent. sulphuric acid. A violet color indicates the presence of rosin.

Test for fillers.—Pure soaps are soluble in hot alcohol; the substances used as fillers are not; therefore the test for fillers is to dissolve the soap in warm alcohol and filter the solutions. If fillers are present, an insoluble residue will remain on the filter paper.

Test for free alkali.—But a few small pieces of soap into a test tube with alcohol and shake the tube until the soap is dissolved, then add a few drops of phenolphthalein. A pink color shows the presence of free alkali.

Test for unsaponified fat.—Put a few thin, small pieces of soap into a test tube with either gasoline, ether, or benzene and shake the tube for some minutes; any fat present will be dissolved by these solvents. Then filter the solution, put some of the filtrate in a watch crystal or Petri dish, and let this stand exposed to the air until the liquid has evaporated. If fat is present, it will remain in the dish.

Scouring Powders

The various scouring powders consist chiefly of soap, soda, and, the majority of them, silica and other rough substances such as are used for soap fillers.

Sapolio consists chiefly of soap and sand. Its value as a detergent is dependent chiefly upon the friction it produces.

These detergents are very efficient cleansers, especially such powders as *Dutch cleanser* and *gold* *dust*, but they should not be used on surfaces that are injured by alkalies or those that are easily abraided.

Scouring Agents Other than Soaps and Alkalies

Scouring agents that owe their cleansing properties to the friction they produce when rubbed upon a surface are: (1) whiting and bonami, both of which are finely pulverized chalk, the latter being pressed into cakes; silica, an oxid of the non-metallic element silicon (the silver polish known as *silicon* contains silica compounds and whiting); (3) bath brick, a form of calcareous earth pressed into the form of a brick; it is used chiefly for cleaning steel knives; (4) rottenstone, known also as Tripoli, after the country in which it was first found, a soft stone used for scouring and polishing metals; (5) powdered emery, a variety of corundum, a metal of extreme hardness; it is used for scouring the harder metals, as steel; (6) carborundum, a carbide of silicon that is even harder than corundum. It is used for the same purposes as emery. It is made by heating coke and sand (which is an oxid of silicon) in an electric furnace.

Bleaching Agents

Nature of bleaching.—Bleaching signifies *making white by removing color or dirt by the action of the sun's rays or by a chemical process.* The chemical process may consist in: (1) The union of the bleaching agent with the coloring or staining matter and the consequent forming of a colorless compound or of a soluble one that can be easily removed. Sulphur dioxid (SO_2) and sulphurous acid[1] (H_2SO_3) (which is made by the chemical union of sulphur dioxid and water) act in this manner. (2) Reduction; *i. e.*, the removal of oxygen from the staining agent thus changing it to a colorless or a soluble substance; sulphurous acid and oxalic acid often remove stains in this way and cleansing pastes and liquids used for metals generally containing reducing agents. (3) The liberation of oxygen and the union of this element with the coloring matter thus oxidizing and, consequently, decomposing it. The oxidizing compounds used as bleaches may themselves contain the oxygen—*e. g.* potassium permanganate ($KMnO_4$) and hydrogen peroxid (H_2O_2)—or they may act indirectly,—*e. g.*, chlorin, which is the principal oxidizing agent acting in this way, has a very strong affinity for hydrogen and unites with the hydrogen of steam, setting free the oxygen which, thereupon, unites with the coloring matter and oxidizes it.

Chlorin also, as it is a very active element, sometimes acts by uniting with substances in the coloring matter and forming colorless or soluble compounds.

Chlorin compounds used for bleaching.—The employment of chlorin in gaseous form for bleaching would be both inconvenient and dangerous; therefore, compounds which will part with their chlorin readily are used instead. One compound that is very generally used both in factories, for the removal of the yellow hue characteristic of freshly woven cotton and linen fabrics, and in laundries, for the removal of dirt and stains, is chlorid of lime, known also as bleaching

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[1] The usual method of obtaining and using sulphurous acid for this purpose is to hang the material, after wetting it, in a closet that can be sealed and burn sulphur in the closet.

powder. This is made by passing chlorin gas over moist slaked lime. As the gas passes over the lime it is absorbed by the latter.

Though the use of a bleach in laundry work is not to be recommended, if used with care, a small amount of bleaching powder in solution in the water in which cotton and linen materials are washed will not injure the fabric, but if solid particles of the powder come in contact with the material or if a concentrated solution is used, or if the material is not well rinsed, so as to be absolutely freed from all trace of the bleach, the fibers of the fabric will be soon rotted. Therefore, to prevent damage to fabrics by the use of chlorid of lime two precautions should be taken: (1) the powder is to be thoroughly dissolved in water before use (about one pound of lime to 12 gallons of water); (2) a substance that will neutralize the action of the bleach should be added to the first rinsing water. Sodium thiosulphate, commonly known as *hyposulphate of soda* is often used for the purpose, for it interacts with the chlorid of lime forming sodium hypochlorite and calcium sulphate both of which compounds will be removed from the material by second rinsing. Chlorin cannot, of course, be used for removing stains from colored material and it cannot be used for bleaching or cleaning silk and wool. See experiments in Chapter XVII.

Sodium hypochlorite, known also as *Labarraque's solution* and *Javelle water*, is another common chlorin bleaching agent. This is somewhat more expensive than the bleaching powder, but it is often preferred to the latter for laundry purposes as it does not react with soap to produce an insoluble calcium soap as calcium chlorid does. The action of sodium hypochlorite should be neutralized by the addition of some such substance as a dilute solution of hydrochloric acid or oxalic acid to the first rinsing water.

For the removal of certain stains (see page 267) sodium chlorid, water, and lemon juice are used. These act as follows: The citric acid of the lemon juice unites with the sodium of the salt, setting free the chlorin, which, thereupon, unites with the hydrogen of the water, setting free the oxygen which unites with the staining agent and oxidizes it. This process, like many others, is hastened by exposing the stain, during the treatment, to the sun's rays.

Absorbents

Material that will absorb readily, such as starch, shredded blotting-paper, and fuller's earth are often used, either alone or in connection with other agents, for the removal of liquid staining matter.

Fuller's earth, which is a form of clay, is so called because used by fullers (certain cloth workers) to absorb the oil or grease with which woolen cloth is sometimes treated in the manufacturing process.

After reading the foregoing pages, it will be realized that various acids and alkalies and, for cotton and linen materials, chlorin are, either alone or as constituents of compounds, the detergents in most common use. It will be interesting, therefore, to test the action of these things on some of the materials used for ordinary household utensils and furnishings and for clothing; viz., metals, marble, porcelain, glass, paint, and cotton, linen, silk, and woolen fabrics, and to consider briefly the source and nature of these materials.

Metals

Before testing the action of cleansing agents on metal, it will be well to consider the nature of the discoloration known as *tarnish*, for this and, especially in the case of kitchen utensils, fat are usually the two most common soiling agents of metals and the ones that are most difficult to remove.

Tarnish.—Tarnish of metals, except that of silver and one form of copper tarnish, is generally due to the union of oxygen with the metal. The majority of tarnishes are thus oxids of the metals. Nearly all metals form oxids when heated, but certain ones, notably iron and lead, do so rapidly at ordinary temperatures if there is moisture present. Under the same conditions, aluminium, zinc, and nickel tarnish slowly. Copper tarnishes very rapidly if heated, even slightly, and the CO_2 of the air, as well as oxygen and moisture, is responsible for the reaction; thus, the common tarnish of copper is more often a carbonate than an oxid of copper. The tarnish of bronze and brass is similar to that of copper. Gold and platinum do not form oxids in the air. Tin and silver do so only at high temperatures, but the latter tarnishes very readily in the presence of sulphur. The usual tarnish of silver is therefore silver sulphid (Ag_2S).

There is an important difference, easily perceived, between the oxid that forms on iron and on other metals; viz., iron oxid (Fe_2O_3), known as *rust*, forms scales or granules which rub off, and thus a fresh surface is constantly exposed to the air and the metal may soon become much eroded. The tarnish that forms on the other metals, on the contrary, occurs as a thin film that adheres to the surface, dulling its luster and causing discoloration, but, being adherent, it prevents air getting to the metal and its oxidation is thus soon stopped; consequently, the metals are not permanently injured. Rust on iron is partly ferric hydroxid [$Fe(OH)_3$], because some moisture, the presence of which is necessary for the oxidation of iron at low temperatures, enters into the reaction.

As will be seen in the descriptions of the metals, those which tarnish very readily, as iron and copper, are often protected from the air by coating them with a metal, lacquer, or other substance that does not unite readily with oxygen.

Requirements of cleansing material for metals.—Since tarnish and fat are the two most common forms of soil on metal, something that will remove them is essential for an efficacious metal detergent. Acids and alkalies will, as previously stated, remove fat, and certain acids, partly by dissolving and partly by reduction, will take off tarnish, but as will be seen in the experiment with metals neither acids nor alkalies can be used with all metals. The hydrocarbons, as kerosene and gasoline remove fat and some oxids. Whiting, silicon, and similar soft, but granular substances, are valuable, for they help to provide the friction necessary to remove soil, and they, when metal is rubbed with soft material, such as chamois, help to produce a thin film over the surface of the metal which reflects light and thus gives luster or polish.

The majority of patented preparations consists of different combinations of substances here mentioned or similar ones, and good preparations contain the ingredients in proportions and combinations that frequent trials have shown to be most efficacious; therefore, though it costs somewhat more to buy the combined ingredients, it is often advisable to do so. Preparations which contain acids should not be used on metals that are injured by acids and those containing alkalies on metals affected by alkalies. The presence of alkalies or acids in cleaning compounds can be determined by the litmus test. Of course, the acids, etc., likely to be present in cleaning preparations would not be as strong as those used in experiment 28, and in such case the effect would not be as pronounced and might be visible only after repeated applications. Rough granular cleansing agents, as emery, should not be used on the softer metals.

Suitable detergents to use for the different metals will be mentioned in the paragraphs describing the metals.

Silicon and whiting are nearly always prepared for use by mixing the powder with water or dilute alcohol, and to get the best results from these and nearly all other kinds of cleansing pastes, the metal should be first washed with hot soapsuds, rinsed, dried, and the paste applied with a soft, clean cloth, allowed to dry, and then removed by rubbing with a soft, clean cloth, or chamois skin.

Experiment 28. Object: To study the effect of acids and alkalies on different metals.

Material required: Small (about ½ inch) pieces of bright and tarnished metal; preferably, aluminium, copper, iron, zinc, and tin; an inorganic acid, preferably hydrochloric; an organic acid such as acetic; sodium carbonate solution 10 per cent.; ammonia water; mercuric chlorid; evaporating dishes or test tubes; an iron stand; forceps.

Procedure: Take three bright and three tarnished pieces of each of the metals and put each piece into a separate evaporating dish or test tube. To a bright and to a tarnished piece of each metal add hydrochloric acid; to another piece of each metal, both bright and tarnished, add acetic acid; to the third set add sodium carbonate solution.

The interaction that takes place when some metals are treated with acids and alkalies is so great that visible results occur immediately; some metals are not acted upon at all by these substances and others are acted upon so slightly that a further test is necessary to see if any of the metal was dissolved. If it was, it will be in solution in the liquid, and the test will consist in adding something to the liquid that will, if the metal is present, combine with it and form a colored compound. It will be seen that there are some metals that, themselves, are not affected by acids and alkalies and that have oxids that are soluble in either or both such liquids.

The test is as follows: After the metals have stood for a short time in the acids and alkalies, remove them with forceps and test the liquids, for iron, aluminium, and copper, by adding ammonia, and for tin by adding mercuric chlorid.

If iron is present, a red color will develop.

If aluminium is present, a white precipitate will form.

If copper is present, a blue color appears.

If tin is present a white precipitate forms that turns black when the liquid is heated.

What is formed by the action of hydrochloric acid on aluminium?

(N. B. Foods containing acids should not be put into utensils that are acted upon by acids.)

Occurrence of the metals.—The metals are usually found in nature in the form of oxids or of salts though a few, notably gold, platinum, and copper, sometimes occur uncombined. The metal salts, as well as some other inorganic substances, are called *minerals*, and minerals from which useful substances can be extracted, especially metals, are called *ores*.

The metals most generally used for household utensils and hospital laboratory appliances are iron, zinc, nickel, tin, copper, aluminium, silver, and platinum.

Varieties of iron.—The iron of commerce is never pure, but contains varying amounts of other elements, especially carbon, phosphorus, silicon, sulphur, and manganese. The percentage of these elements, especially copper, present in iron changes its nature considerably, and the amount of carbon and the form in which it is present gives rise to three distinct varieties of iron which are known respectively as *cast iron*, *wrought iron*, and *steel*.

Cast iron and **pig iron** are names given to the metal as it comes from the blast furnace. This contains more carbon than the other varieties. It is hard and brittle and, therefore, cannot be forged or welded into shape, but must be melted and cast in molds. It has, compared to wrought iron, a low melting point; *i. e.*, 1100°. Cast iron retains its heat better than other forms of the metal. It becomes, if frequently heated to redness and cooled suddenly, very brittle, as is demonstrated by the cracking of stove covers. If red-hot iron is covered, so that cooling takes place slowly, it will not become so brittle.

Stoves, furnaces, and radiator pipes are made of cast iron.

Wrought iron is made by burning out some of the carbon, phosphorus, etc., from cast iron; thus it is a purer form of iron than the latter. It has a higher melting point than cast iron (1600°), but is not as hard and, consequently, it can be drawn out into wire, rolled into sheets, and wrought into various forms.

Steel, like wrought iron, is made by extracting most of the other elements from cast iron, but the process is carried out in a different way. Since the means of making steel at a reasonable price have been provided, steel has taken the place of wrought iron for many purposes, for though, like the latter, it is, when heated, quite malleable and can thus be forged into shape, rolled into sheets, etc., it can be also cast in molds and it is much harder and stronger than wrought iron. Another exceedingly valuable property of steel is that it can be easily tempered; *i. e.*, made harder or softer as required. For example, if steel is heated to about 170° C. and cooled suddenly, as by plunging it into very cold water, it becomes exceedingly hard, but, if it is heated to about 235° C. and cooled slowly, it will be so pliable that watch-springs and fine wires can be made from it. The difference in temper is largely caused by differences produced in the nature of internal combination of the carbon by the different temperatures and method of cooling. In the softer steel, the greater part of the carbon is present in the form of graphite; in the harder steel, the carbon is in the form of a carbide.

Galvanized iron is the name given to iron covered with zinc. It was so named after Galvani, the discoverer of the electric current, because, formerly, the zinc coating was obtained by means of electrolysis. At the present time, however, the more common method is to dip the iron into molten zinc and then pass it between rollers. As seen in Experiment 28, when acids and zinc come in contact, the zinc is dissolved and poisonous zinc salts are formed; therefore, galvanized iron cannot be used for culinary utensils, but as zinc seems to prevent iron rusting better than the majority of substances used for the purpose, galvanized iron is in great demand for out-of-door use.

The **enamel ware utensils** in such common use for culinary utensils are made of iron covered with some form of glaze. The glaze is usually obtained by dipping the iron utensil into a molten solution of fusible silicate (see page 253) and drying and hardening it in a furnace.

Enamel is not affected by acids nor weak alkaline solutions; it is light in weight and easily cleaned. These qualities all enhance its value for culinary uses, but sudden cooling after heating, over-heating, as by leaving the utensil over the flame when it does not contain anything that will conduct the heat away, blows, scraping to remove matter that has dried upon the utensil, cleaning with strong alkalies, all tend to crack the enamel glaze and when this happens, the exposed iron rusts more easily than that which has not been so protected.

The characteristics of iron that allow of its being used for the various household purposes to which it is put are: (1) its infusibility (*i. e.*, it can be melted only at extremely high temperatures); (2) its light weight as compared with the majority of metals; (3) the readiness with which it radiates heat. Two characteristics of iron that are defects as far as such purposes are concerned are the manner in which it reacts with acids and the ease with which it unites with oxygen and thus becomes *rusty*. Because of the chemical action that occurs when acids and iron come in contact, foods that contain acids should not be cooked in iron utensils unless they are protected with enamel or other substance that is not affected by acids, and only in such case should acids be used for cleaning ironware. To prevent iron rusting, it is necessary to keep it dry and clean. Highly polished steel or a well-blackened stove will not rust nearly as quickly as soiled or dull metal.

Kerosene is one of the best substances to use to remove rust from iron. Stove blacking, the principal constituent of which is graphite, a form of carbon, is used for polishing stoves, stove-pipes, and the like. As steel is very hard, such substances as emery, carborundum, and bath brick can be used for cleaning and polishing it. Alkaline solutions can be used for scouring galvanized iron, but only dilute ones should be used for enamel ware. Lemon juice, vinegar, and dilute oxalic acid can be used for removing stains from enamel ware, but not from galvanized iron. A thin paste of whiting and kerosene is excellent for cleaning zinc and, consequently, galvanized iron.

To prevent utensils, instruments, etc., made of any variety of iron rusting where they are stored, they are covered with some substance—as oil or vaseline—that will protect them from the air and, if they are not to be used for some time, they are wrapped in soft paper or other covering that will prevent the oil being removed or dust collecting upon it.

Nickel is a somewhat rare metal element that is used principally as an alloy with, or as a covering for, other metals. It can be deposited upon the surface of metals by electrolysis. It is often used in this way with iron, for it will take a high polish and it will protect the iron from rust, for the tarnish that forms upon nickel, even at high temperatures, is very superficial. A paste made of whiting and either dilute ammonia, alcohol, or kerosene is suitable for cleaning. Very badly stained nickel can be cleaned by washing with: (1) a solution of one part sulphuric acid and 50 parts alcohol; (2) with alcohol.

Pure tin, known as block tin, is a soft white metal element with a silver-like appearance. It melts at a temperature of 228° C. It does not form oxids readily and is therefore often used to cover iron and other metals which do. The tin generally used for utensils, roofing, and similar purposes is known as *tin plate* and consists of iron covered with tin, the tin coating being applied by the same methods as is zinc in the manufacture of galvanized iron. The low melting point of tin is its objectionable quality for cooking purposes. If an empty tin utensil or one lined with tin is left over the flame even for a few minutes, the tin will be melted.

As seen by Experiment 28, pure tin is not acted upon by weak acids or alkalies, but it is by strong solutions; thus it can be used for cooking fruit and the like and it can be washed with weak sodium carbonate solution; but only weak solutions of either acids or alkalies should be used for cleaning it. Tin containing impurities is sometimes acted upon by acids and poisonous salts formed by the reaction. A soft paste of whiting and kerosene, plus rubbing, will both clean and polish tin.

Copper is a strong, heavy metal, but, when highly heated, it becomes soft and malleable. It melts at

1084°. Next to silver, it is the best known conductor of electricity; therefore, as it is cheaper than silver, it is much used for electrical purposes. Copper is one of the best heat conductors and it does not radiate heat readily, but retains it much better than the majority of metals. These qualities make it a very good material for cooking utensils such as pots and saucepans, especially for the cooking of large quantities of food and when the utensils are subjected to hard usage, for copper is practically indestrucible. Unfortunately, it is heavy and hard to keep clean, and though, as shown in Experiment 28, bright copper is not acted upon by acids or alkalies, copper oxid and copper carbonate, popularly known as *verdigris*, are soluble in acids and they interact with acids to produce poisonous salts. For this reason, copper cooking utensils, unless kept very clean, are dangerous to use. To obviate this danger, copper utensils are often lined with tin, but, as already stated, tin is likely to be melted or removed by other means and the exposed copper tarnishes very readily and is exceedingly difficult to keep clean. Copper utensils can, however, be retinned.

Copper is extensively used for making alloys, *i. e.*, combinations of metals. Some of the more important of these are: Brass, which consists of copper and zinc; bronze, which is copper, tin, and zinc; German silver, which is copper, zinc, and nickel; gun metal, which is copper and tin.

The outer surfaces of copper appliances are sometimes covered with lacquer, consisting, usually, of shellac, and sometimes a coloring substance dissolved in alcohol. Lacquer is spotted by water, and therefore a lacquered surface should not be washed, but should be kept clean by rubbing with a soft, clean, dry cloth.

Both acids and alkalies can be used for cleaning copper and solutions of both oxalic acid and ammonia either alone or in combination with such substances as whiting, silicon, and kerosene are much used; as are also putty powder, rottenstone, and various patented preparations. A hot solution of sodium chlorid in vinegar (1 part of salt to 16 of vinegar) is very effectual in removing stains from both copper and brass.

Aluminium is a tin-white metal of very light weight. On account of this latter quality and its nice appearance, aluminium is much used for cooking utensils. It has, however some disadvantages for cooking purposes, viz., it radiates heat so rapidly that food cooked in it takes longer to become hot than when cooked in utensils made of other metals, and it will not remain hot in aluminium utensils for any length of time after the latter are removed from the fire. Also aluminium is injured by both organic and inorganic acids, by alkalies and by salt solutions; therefore neither fruit nor other acid food nor salty food should be cooked in aluminium utensils, nor should milk be kept in them, in case it should sour.

Aluminium is best cleaned with neutral soap and water. A paste made of whiting and very dilute ammonia can be used when necessary. Two patented preparations, the *Universal Metal Polisher* and *Putz Pomade*, have been found very effective in removing from operating-room utensils stains made with bichlorid and iodin and have not injured the metal. As aluminium is so easily injured by the common constituents of metal-cleaning preparations, caution must be observed in using those the effect of which is unknown.

Silver is a heavy, rather soft, white metal, that can be highly polished. It is not affected by organic acids and does not oxidize at ordinary temperatures, but the presence of even very minute traces of sulphur, such as are usually present in the air, especially in the vicinity of places where gas is burned, will cause the silver to tarnish, for the sulphur will unite with it and form a thin film of silver sulphid (Ag_2S) over the surface. It is because of the sulphur they contain that eggs discolor silver spoons and, for the same reason, woolen fabrics and material for the bleaching or dying of which sulphur has been used may cause a similar discoloration. It is therefore well, when storing silver articles, to coat them with vaseline and wrap them in soft unbleached muslin or in tissue paper.

Whiting, rouge, and silicon are all much used for removing tarnish from, and for polishing, silver. An effective substitute consists in boiling the articles to be cleaned in a solution of sodium chlorid and bicarbonate of soda (one tablespoon of each of these salts to a quart of water) in a patented apparatus known as a *silver-clean pan.* The salts are decomposed and the soda unites with the sulphur forming a salt that is soluble in the water. After the silver has been boiled for three minutes, it is removed from the solution, rinsed in water, and dried with a clean, soft cloth. In hotels and other places where a large number of articles must be cleaned daily, the polishing is done by a polishing machine run by electricity, and sometimes especially large articles are cleaned by dipping them in hydrocyanic acid. The acid unites with the sulphur of the tarnish forming a salt that is soluble in the acid. This is a very quick and effective method, but it must be carried out with caution and the fingers must not come in contact with the acid for it is exceedingly poisonous.

Lead is a heavy, lustrous metal that owes much of its usefulness to its softness and to the ease with which it is melted. It tarnishes readily and is acted upon by the majority of acids, except hydrochloric and sulphuric, forming salts that are injurious to health for, though pure lead is not poisonous, many of its salts and other compounds are. A common form of slow lead poisoning is that which painters often contract. The usual cause of this is taking their food in their hands while they are soiled with paint containing lead compounds.

The majority of lead compounds are insoluble in pure water, but soluble in water containing carbonates and sulphates. Therefore, water containing such salts, after standing in lead pipes, may contain injurious lead compounds, but the danger, it is thought, has been overestimated, since, if such salts are present in water, they soon form a coating over the interior of the pipes. However, as lead pipes are much used in houses, it is a wise precaution, when the plumbing is new and when water has not been drawn for some time, as, for instance, when a house has not been lived in for several weeks, to let water that is to be used for drinking or cooking flow for some minutes before use, so as to empty the pipes of that which has stood in them.

An alloy of lead and tin, known as *solder*, that is easily melted, is much used for cementing severed metal surfaces together.

Pewter is an alloy of tin and lead and therefore acid food should not be put into pewter utensils nor should they be cleaned with acid-containing substances.

Platinum is a grayish-white metal of high luster. It is a good conductor of electricity. It is very malleable and ductile, but fusible only at extremely high temperatures. It is very inert (*i. e.*, it does not combine readily with other elements) and therefore it does not tarnish and is not attacked by many acids nor alkaline substances. These qualities make platinum of great value for utensils used in electrical, chemical, and bacteriological work. Platinum is acted upon by nitric acid and by chlorin.

Marble

Experiment 29. Object, to study the effect of acids and alkalies on marble.

Material required: Powdered marble or calcium carbonate ($CaCO_3$), hydrochloric acid, lemon juice, ammonia water, sodium carbonate solution 10 per cent., 6 test tubes.

Procedure: Put a little calcium carbonate into each of the six test tubes and add to (1), hydrochloric acid; to (2), lemon juice, to (3) ammonia, to (4) sodium carbonate solution; to (5), water. Let the five test tubes stand. Watch results. To tube (6) add some hydrochloric acid and, when action is well established, some ammonia water.

It will be seen as the result of this experiment that marble is insoluble in water and alkaline solutions, but that it interacts with, and is dissolved by, acids, especially mineral acids. Also that an alkali, by neutralizing the acid, stops the reaction. Since acid has this effect upon marble, if lemon juice, or other acid, is spilled upon it, the surface of the latter will be roughened unless the action of acid is at once neutralized by the addition of ammonia or other alkaline substance. Marble is best cleaned with hot soapsuds or alkaline solutions. Many stains can be removed by rubbing them with NaCl or kerosene.

Marble was formerly much used in kitchens and lavatories, but the ease with which it was affected by acids led to the exclusion of its use for many purposes as soon as the means of making suitable porcelains and cements were discovered. There are now a number of cements and cement-like substances used as tiling, and in other form, for floors, table covers, sinks, etc., that are not affected by either acids, alkalies, or heat and that do not absorb grease—most important qualities to be considered in judging the value of such material for kitchen use. Tiling tinted gray, buff, or other dull hue is much better than white for floors and table tops, because (1) white shows every mark and (2), if the kitchen is as light as it should be, white causes a glare that is very trying for the eyes.

Glass

Composition.—Glass consists of mixtures of different bases and various silicates (*i. e.*, salts of silicic acid, an acid that is obtained from silicon, a form of sand).

Crown glass, which is used for windows, consists of silicates of soda and lime; plate glass, also used for windows and for mirrors, consists of silicates of soda, lime, and potassium; flint glass, which, as it has a high refractive power, is much used for optical purposes, consists of potash and lead silicates; Bohemian glass which, being very infusible, is much used for chemical purposes, is a silicate of potash and lime; the cheaper glasses, such as are used for bottles, consist of the same ingredients as Bohemian glass, but they contain impurities.

Glass is made by fusing silica—in the form of sand or quartz—with the desired bases, at an exceedingly high temperature, for a number of hours. The molten mass, when clear, is made into the desired shape, usually by blowing or pouring it into molds. It is then cooled rapidly, in order that it may retain its transparency, and afterward it is annealed, as this renders it less brittle. Annealing consists in heating the glass to a temperature just below its melting point and then allowing it to cool slowly.

Opaque glass is made by adding some finely divided infusible matter to the molten glass before it is shaped. Colored glass is made by the addition of coloring matter—usually a mineral oxid—to the molten glass; *e. g.*, a blue color is obtained by the use of an oxid of cobalt; a red color by the use of copper or gold oxid.

Porcelain and Chinaware

Porcelain and chinaware consist of different silicates of aluminium. These occur in certain clays. For the finer porcelains, all impurities are removed. In the making of dishes, etc., the clay is moistened with water, molded into shape on a wheel, and heated in a furnace to a strong red heat. Large portions of the silicates in the clay used are infusible, but some are fusible, and these, while the articles are in the furnace, melt, and this molten matter, spreading through the mass, cements the infusible substances together.

When the articles are removed from the furnace, they are porous, and in order to render them impervious to liquids, they must be glazed. This is done by dipping them in a mixture of molten, fusible silicates and returning them to the furnace that the *glaze* may harden. The glaze is practically the same as glass.

Cleaning of glass and porcelain.—If hot alkaline solutions are used on glass and porcelain, some of the alkali combines with the silicates, and repeated treatment of this kind will diminish the luster of the glass or glaze; therefore, only neutral soaps or soap powders should be used for washing good glass and porcelain ware. This is true also of glazed or enameled sinks and utensils. Kerosene, benzene, and whiting can be used, and are usually very effectual for removing stains that soap will not eradicate. Glass and glaze are rather easily scratched, so that nothing rougher than whiting should be used. For cleaning windows, mirrors, and the like, a thin paste of whiting and water with, sometimes, a little soap is rubbed over the surface of the glass; the water is allowed to evaporate, and, afterwards, the dried whiting is rubbed off with a soft, clean, lintless cloth.

Paints and Varnishes

These substances serve two purposes: (1) they, except water paints, protect the surfaces they cover from the influence of air, moisture, etc.; (2) they are used for decoration.

Nature of paints.—Paints are classed as water and oil paints and the water paints used in masonry are classed as whitewashes and kalsomines.

Whitewashes are usually either mixtures of water and lime, boiled soap and lime, or skimmed milk and lime, and, if colored, coloring matter. When a whitewash is spread upon a surface it loses water by evaporation and absorbs CO_2, thereby becoming changed to chalk ($CaCO_3$).

Kalsomine consists of a finely ground chalk and coloring matter suspended in a warm solution of glue. It is applied while hot, and as it cools, it loses water by evaporation and becomes dry and hard.

Since both whitewashes and kalsomines become dry by the loss of water, they will be dissolved by the addition of water. They cannot therefore be used on the exterior of buildings nor can surfaces covered with them be washed. Water paints, however, are cheaper than oil paints and are therefore often used for coloring walls of new houses tor, as the plaster is apt to crack during the first months following structure, due to shrinking of the wooden rafters, etc., the walls generally need to be repainted or papered in a year or two.

Oil paints consist of coloring matter (which usually is composed of ground minerals), carbonate of lead (known also as white lead), or zinc white, or both held in suspension in what is called a *drying oil*—linseed or poppy oils being the ones most commonly used.

The carbonate of lead and zinc white are bases and, when mixed with oil, they form a soap which, as sufficient oil is used for the purpose, is held in solution in the latter.

The oils used are called *drying oils* because, especially when spread out in a thin film, they unite with oxygen and, when oxidized, become dry and hard.

Substances that will hasten the absorption of oxygen—as compounds of manganese—are often added to paints to hasten the oxidation and, consequently, the drying. When used for this purpose, such substances are called *driers*. Sometimes, instead of using driers which act in this manner, matter that will readily part with its oxygen—as lead oxid—is used. Spirits of turpentine, called *turps* by painters, is sometimes added to paint to thin it or, especially in the first coat of paint applied to wood or plaster, to hasten the absorption by the surface upon which it is spread.

Methods of cleaning painted surfaces.—As water paints are soluble in all liquids, including water, surfaces covered with them cannot be washed, therefore dusting with a soft, dry cloth is the best means of cleaning them. Oil paints are insoluble in water, and surfaces covered with them can therefore be washed with warm water and neutral soap, or, with white paint, bonami can be used. (N. B.—Bonami and whiting should not be used on colored paints.) Alkalies unite with some coloring substances and thus cheap scouring soaps and other alkaline detergents should not be used. Neither should coarse, hard, cleansing agents be employed, for paint is easily scratched.

Nature of varnishes.—Varnishes are usually classed as spirit varnishes and oil varnishes.

Spirit varnishes consist of gum resin, such as shellac, dissolved in alcohol; oil varnishes of fossil resin, such as amber, dissolved in hot linseed oil. When either form of varnish is spread over a surface, the solvent—the alcohol or oil—evaporates and the resin solidifies.

Shellac is rather soft and is spotted by water, therefore spirit varnishes are not suitable for surfaces exposed to weather or to rough usage.

Cleaning varnished surfaces.—Good oil varnish can be wiped with a soft, damp cloth, but the surface must be wiped dry, and water should not be allowed to get into seams of the woodwork. If the varnish is inferior, however, the use of a damp cloth will soon cause yellow or white spots to appear. A weak solution of good Castile or other neutral soap can be used to remove grease stains if the soap is washed off with a cloth dampened in clean warm—never hot—water. After such treatment, a varnished surface should be rubbed with a soft cloth moistened with a reliable varnish polish. Too much polish must not be used and the surface to which it is applied must be well rubbed, for such polishes nearly all contain some oily constituents and, if the surface is left wet, dust may collect and be difficult to remove.

Both paint and varnish are to some extent soluble in alcohol, therefore if the latter is spilled on furniture it will cause spotting, unless its action is neutralized by pouring oil over it. This should be done before wiping off the alcohol. White spots on furniture, due to exposure to heat, may be removed by rubbing them with spirits of camphor, or equal parts of linseed oil and alcohol.

Wax

Waxing is a very common way of treating floors. Before wax is applied to a new floor, the latter should be given two coats of a mixture of equal parts of pure linseed oil and turpentine with sufficient japan to make the mixture harden overnight. This brings out the grain of the wood and prevents subsequent grease spots showing. The floor waxes commonly sold are mixtures of beeswax, paraffin, and turpentine. Before applying the wax, it is, if necessary, thinned to the consistency of a paste with turpentine. It is most important not to put the wax on in a thick layer, for when this is done the wax takes a long time to dry thoroughly and it collects dust. After the wax is applied, it is left for about twelve hours to harden, and then a little more paste is rubbed over the floor with a soft cloth. The turpentine in the paste dissolves any surplus wax that may be on the floor and makes the coating more uniform. When the wax is perfectly dry, the floor is polished with a weighted brush and the rubbing should be done with the grain of the wood.

For floors likely to be spotted with water, it has been recommended that they be saturated with paraffin by covering the floor with powdered paraffin, such as is scattered over the floor when preparing for dancing, and melting the paraffin with hot irons or to have the floor waxed in the usual manner and sprinkle powdered paraffin over it, and then, as in the ballroom, allow the feet of those passing over it to force the paraffin into the wood. This treatment gives the floor a hard, impervious surface, which will not be easily spotted with water.

Cleaning waxed floors.—After a floor is properly waxed, the only treatment that it usually requires for some time is an occasional polish with a weighted brush. When the surface grows dull another thin application of wax can be made in the same manner as the previous one. If water is spilled, after it is dried, the subsequent spots can be removed by rubbing them with wax and then with a warmed woolen cloth and a weighted brush. Oil should never be put on a waxed floor as it causes the wax to become dark and spotted. Ink stains can be usually removed by putting a little solution of oxalic acid over the spots and allowing it to remain until it dries; the oxalic acid unites with the iron in the ink to form an oxalate of iron, which can be easily washed off. The subsequent white spots can be obliterated by rubbing a little wax into them. There are some inks which are made from the aniline colors and do not contain iron, and to remove these and, also, grease spots, sodium carbonate can be used (except on oak floors, which are stained brown by alkalies). White spots caused by the soda must be treated as those due to other causes.

CHAPTER XVII

TEXTILES

Classification and Origin of Textiles—Effect of Detergents on Different Kinds of Textiles—Methods of Removing Stains—Textile Tests—Printing and Dyeing—Nature of Dyes—Fading of Colors.

THE various kinds of materials used for clothing, upholstery, and the like are classed as textiles.

Source of textiles.—Textiles are made by spinning, weaving, and otherwise treating fibers which are obtained from animals, plants, and, in a few cases, inorganic matter.

From fibers of animal origin are made wool, hair, and silk; the principal vegetable fibers are cotton, flax, jute, and hemp; the chief mineral fibers are asbestos, and the spun glass, silver, and other minerals classed as *tinsels*.

Vegetable fibers consist mainly of cellulose. The cotton fibers from which cotton materials are made are procured from the cotton plant. Flax fibers, obtained from the plant of the same name, are, when woven, known as linen.

The quality and nature of woolen materials depend both upon their source and the processes to which they are subjected in the course of manufacture. The wool of different animals varies greatly in strength, length, fineness, and softness and so does that of similar species of animals that are reared under different conditions and in different climates. Also, the quality of wool taken from different parts of an animal varies very considerably.

The chief difference between wool and hair is the greater fineness of the fiber of the former and its highly serrated surface.

The fibers from which silk material is spun are produced by silk worms, from liquid which they secrete, in much the same way as spiders spin their webs.

Both silk and wool contain protein substances; this can be proved by Experiment 30.

Experiment 30. Procedure: Dip some pieces of white (1) all wool material, (2) mixed wool and cotton material, (3) silk, (4) linen, (5) cotton, into some Millon reagent (see page 284), and then allow them to dry. The wool and silk will assume the red color which always indicates the presence of protein (see page 284). In the mixed wool and cotton material the cotton threads can be easily discerned, for they remain uncolored. This is one of the tests used to discover if a material is all wool or a mixture of wool and plant fibers.

Experiment 31. Object: To see if there is sulphur present in the different textiles.

Procedure: Burn a little of each kind of textile in a flame, notice the difference in the odor from the wool. Hold pieces of paper moistened with lead acetate solution in the smoke. Hydrogen sulphid (H_2S) reacts with lead acetate to form lead sulphid (PbS) which is black.

Wool and hair are the only textiles that contain sulphur as a part of their actual composition, but sulphur may be present in new materials of other kinds for the bleaching or dyeing of which sulphur has been used.

Ravel out threads of each kind of textile and examine them under the microscope.

The cotton fibers, though orignally tubular, become flattened during the manufacture of the textile and, under the microscope, show a characteristic twist with the ends gradually tapering to a point.

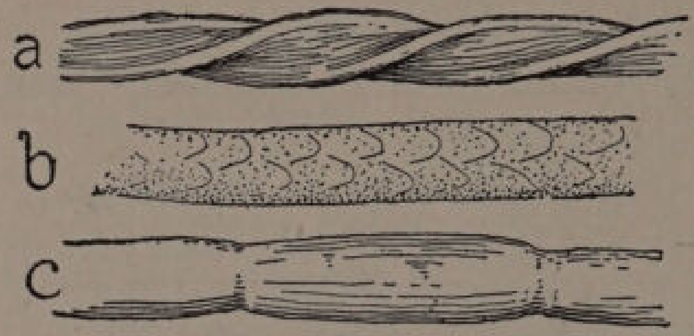


FIG. 60. (a) Cotton fiber. (b) Wool fiber. (c) Linen fiber. As they appear when viewed through the microscope.

Linen fibers, it will be seen, resemble cotton, but they are straighter and more tubular and they have slight serrated notches or joints along their walls which serve to hold the fibers together. These teeth in the notches are not so long as those in the wool fiber and thus, though linen shrinks, it does not do so to such an extent as wool.

The wool fibers, it will be noticed, are marked by transverse serrated divisions. When woolen material is rubbed, or subjected to hot water or strong alkalies, these teeth-like processes become curled, knotted, or tangled together, and this causes shrinking of the material. If cotton fibers are mixed with the wool in a textile, there will be less shrinking.

Silk fibers are perfectly smooth and when rubbed they slide over each other. This sometimes causes a slight shrinking in the width of silk textiles.

Mercerized textiles.—Mercerized materials were so called after a man named Mercer, who introduced the process. This consists in treating the fabrics with caustic soda or sulphuric acid; the former being used for cottons and the latter for wools. This makes the fibers heavier and flatter and gives them a silky appearance and feel.

Artificial silks.—Various kinds of artificial silks are made from cellulose by different processes. They lack the tensile strength and elasticity of silk and thus do not wear as well.

Asbestos is a silicate of magnesium and lime. It can be spun into a fine thread and then woven into cloth. It is incombustible.

Experiment 32. Object: To test the action of acids, alkalies, and chlorid of lime on different textiles.

Procedure: Put small pieces of the different textiles into separate beakers containing: (1) concentrated oxalic acid; (2) oxalic acid 1%; (3) concentrated hydrochloric acid; (4) hydrochloric acid 5%; (5) hot sodium hydroxid solution 10%; (6) cold sodium hydroxid solution 10%; (7) hot sodium carbonate solution 10%; (8) cold sodium carbonate solution 10%; (9) a hot saturated solution of borax; (10) a dilute solution of borax; (11) a hot dilute solution of ammonia; (12) a hot saturated solution of chlorid of lime; (13) a dilute solution (10%) of chlorid of lime; (14) dry chlorid of lime. Use colored pieces of material for the lime tests.

Let the pieces of material remain in the solutions or, in 14, covered with the lime for half an hour, then remove them and cut each piece in half (be careful not to get those from the different solutions mixed); rinse one set of pieces thoroughly in hot water and place them, and also those which were not rinsed, where they will dry. Note and record any changes that have occurred in the fabrics; pull the pieces of material and see if their fibers have been weakened.

State which of the detergents can be used for each kind of material and in what strength.

Why was the color not affected by the dry lime?

Cotton is not much affected by alkaline solutions, if the material is well rinsed, otherwise the alkali becomes concentrated as the water evaporates during drying and, as it acts for some time, it weakens the fibers of the material.

Cold dilute acid solutions do not injure cotton, if the material is well rinsed; strong solutions either cold or hot do, hot ones being more injurious than cold.

The action of detergents and bleaches on linen is about the same as on cotton, but linen is more easily affected by alkalies.

With the exception of nitric acid, dilute acid solutions do not injure wool. Why does nitric acid affect it? (See protein test, page 284.) A warm solution of sodium or potassium hydroxid will dissolve wool; the carbonates of these alkalies have not such an injurious effect, but they make it hard and harsh. Even weak solutions of chlorid bleaches injure wool.

Experiment 33. Object: To test the effect of dry heat on wool. Pass a hot iron over a small piece of woolen material several times and note the result. Cover a similar piece of material with a piece of damp muslin, and pass the iron over it. Compare this piece with the other and with a piece that has not been pressed.

The protein matter of wool contains considerable water which heat causes to evaporate, thereby causing loss of luster and strength of the wool fibers. Why was the piece of material covered with the damp cloth not affected in the same way as that pressed without it?

Silk is affected by the same substances as wool and in the same way except that it is not injured by moderately high degrees of heat and it is injured by acids.

Some Practical Applications of Knowledge of Effect of Acids, Alkalies, etc., on Textiles

As the result of the foregoing experiments, it can be seen that weak soda solutions (the strength depending upon the hardness of the water, see page 220) can be used in washing cotton and linen materials, if the fabric is well rinsed, this care being especially important in the care of linen. Only weak borax solutions and neutral soaps should be used for woolen and silk materials. The reasons why the addition of an alkali to water used for laundry purposes is advisable were discussed in connection with hard water.

Though it is better not to use bleaching agents in the laundry, weak solutions of such bleaches as chlorid of lime and Javelle water can be used for bleaching cotton and linen, if the precautions mentioned on page 216 are taken, but neither of these substances should be used for wool or silk, or any kind of colored material.

Woolen and silk materials should not be rubbed hard. Woolen material must not be exposed to a high temperature, therefore, it should be washed in warm, not hot, water and dried at a low temperature; it should not be dried in the driers. When pressing woolen goods, a damp cloth should be put between the material and the iron.

Bluing

Bluing is used instead of, or in addition to, bleaching, to renew the white color of fabrics that have acquired a yellowish hue. That it will do this is due to the fact that blue and yellow are complementary colors and when combined in proper proportions are seen as white.

The blues in most common use are: (1) Ultramarine, a prepared compound of the elements aluminium, sodium, silicon, sulphur, and oxygen; (2) various coal-tar products as indigo, indigo-carmine, and the alkali blues; (3) Prussian blue, a compound of iron, carbon, and nitrogen. Prussian blue is considerably cheaper than the others, but it is decomposed by soaps and alkalies and if any of these substances are left in the clothing, or when the goods are next washed, the blue will be decomposed and ferric hydroxid [$Fe(OH)_3$] produced. This will cause yellow stains in the fabric.

Experiment 34. Object: To discover if iron is present in blue.

Procedure: Make a strong solution,—about a teaspoonful of liquid Prussian blue to a cup of water,—have two strips of white muslin cut from the same piece of material, dip one strip in the bluing water, then dry it by pressing it with a hot iron. Compare the muslin thus treated with the other strip. If there was iron in the blue the blued strip will have yellow stains. To prove that this staining is due to ferric hydroxid, use a test for iron. A common test is: to pour a drop of (1) pure hydrochloric acid, (2) yellow prussiate of potash over the stain. If this has been caused by iron, a blue color will develop.

The Removal of Stains from Textiles

Some important things to remember in connection with the removal of stains from woven fabrics are:

(1) If the staining agent is allowed to become dry, it soaks into, and hardens in, the fibers, and it is then more difficult to remove.

(2) Unless the staining agent is soluble in hot water, it will, probably, if put into hot water, be hardened by the heat.

(3) When strong acids or alkalies are used to remove stains, they must be neutralized and the material thoroughly rinsed, or it will be destroyed.

(4) When the detergent used acts by dissolving the staining matter (as when ether is used to dissolve fat stains), some absorbent material, as a loose pad of soft muslin, should be placed under the stained material so that the matter will be absorbed as quickly as it is dissolved, otherwise, it will spread and leave a mark.

(5) When removing stains from colored material, always try the reagent on a seam or other part that will not show; for, as can be easily appreciated from what is said in the section on dyes, colors are affected in different ways by the various reagents.

(6) Exposure to the sun's rays hastens such chemical reactions as that which occurs when salt and lemon juice are used to remove a stain.

(7) The sun is a powerful bleaching agent, consequently the whiteness of white fabrics will be increased by exposure to its rays, and the tint of colored material may be injured; the reason for this was given in Chapter VI.

Methods of Removing Stains

Balsam of Peru stains.—Soak and wash the stained part in alcohol.

Bichlorid stains.—Soak the stained part in a ¼ per cent. solution of Javelle water for twelve hours. Then soak and wash it thoroughly in hot water.

Blood stains.—If blood cannot be washed out with soap and tepid water, soak the stain in peroxid of hydrogen, or wash it in ammonia water.

Coffee and chocolate stains.—These can be usually removed by washing them with hot water; if not, place the stain over a bowl containing boiling water, sprinkle the stain with borax, and wash it with hot water.

Fruit stains.—These can be often removed in the same way as coffee; if this fails, use either of the following methods: (1) Rub the stain with a concentrated solution of oxalic acid; then neutralize the acid with ammonia and wash as much of the material as was wet with the reagents in hot water. (2) Immerse the stained part of the fabric in equal parts of boiling and Javelle water and, after a few minutes, wash it thoroughly in boiling water. Javelle water must not be used for silk or wool or colored material.

Glue stains.—Apply vinegar to the spot with a piece of soft muslin until the stain is removed.

Grass stains.—Wash the stained part of the material in either fels naphtha soap or ammonia and water, ether, or alcohol. If no one of these is effectual, spread a paste made of fels naphtha soap and bicarbonate of soda over the stain and allow it to remain for several hours, and then wash it off with hot water.

Grease stains.—Wash the stained part with either vinegar, hot water and soap, hot soda water, hot alcohol, ether, benzene, carbona, or gasolene. Remember the precautions given in section 6, page 267.

Ink stains.—Stylographic and red inks can be generally removed by washing with soap and tepid water, especially if the washing is done while the ink is still wet. Stains made with other inks can be usually removed with lemon juice, salt and water, or with oxalic acid. To use lemon juice, soak the stained part in warm water until it is thoroughly moistened, then rub and cover the stain with salt and lemon juice and place the fabric where the sun will shine upon the stained portion. Make fresh applications of water, lemon juice, and salt when required. It may be necessary to continue the treatment for two or three hours, then wash the part with (1) ammonia, (2) hot water. Use oxalic acid as directed under fruit stains.

Iodin stains.—Soak the stain in either ammonia, ether, or chloroform.

Iron rust.—Use either lemon juice, salt and water, or oxalic acid as directed for ink stains.

Kerosene.—Soak the stain in hot water and then cover it with a thick layer of moistened fuller's earth. Let this remain for several hours and then wash it off with hot water.

Meat-juice stains.—Wash the stain with either peroxid or hydrogen or else with (1) cold water, (2) warm water and soapsuds.

Medicines.—Either alcohol, ether, or ammonia will usually remove stains made with drugs. For drugs containing iron use the same reagents as for iron rust.

Mildew.—Exposure to sunlight followed by hard brushing will sometimes remove mildew. If this fails, soak, the stained part in lemon juice until it is thoroughly saturated and then place it in the sunlight. If this is not effectual, cover the stain with a paste made of one tablespoonful of powdered starch, one teaspoonful of salt, and the juice of one lemon. Allow this to remain for twenty-four hours. Repeat the application if necessary. Mildew is a species of mold.

Paint.—Rub the stain with benzene, turpentine, chloroform, or naphtha. The two last named are best for delicate colors. Remember the precaution mentioned in sections 4 and 5, page 267.

Perspiration.—Wash the discolored part of the material with soapsuds and, while still covered with the suds, place the garment in the sunlight.

Scorch.—Moisten the scorched material and hang it in the sunlight.

Tea.—See coffee.

Varnish.—Make repeated applications of alcohol, turpentine, chloroform, or vinegar. Chloroform can be used with the majority of colors except blue; use vinegar for blue.

Textile Tests

Experiment 35. Object: To distinguish cotton from linen. Stain a small piece of the material to be tested with fuchine and then dip it into ammonia. Linen fibers will remain red, but cotton ones will part with the dye when wet with the ammonia.

Experiment 36. Object: To distinguish cotton and linen fibers from wool.

Procedure: Boil a small piece of the material to be tested in a 5% solution of sodium hydroxid. Dip another piece into a 5% sulphuric acid solution and put the pieces where they will dry quickly. The cotton fibers are not affected by the alkali and the wool fibers are. The wool fibers are not affected by the acid and the cotton fibers are carbonized.

The Millon test (see Experiment 30) is another method of distinguishing between animal and plant fibers.

Experiment 37. Object: To distinguish silk fibers from wool.

Procedure: Place the fabric in cold concentrated hydrochloric acid. Silk fibers will dissolve; wool will merely swell.

Dressing of Cotton and Linen Fabrics

An important thing to consider in deciding the worth of cotton and linen materials is the amount of dressing substances that they contain.

By dressing is meant such substances as starch, gum, dextrin, glycerin, and various mineral derivatives that are used in order to, in some way, improve the appearance of the material. When the dressing is applied to the warp, it is usually called *sizing;* when it is applied to the woven material, it is termed *finishing.* Some dressing is beneficial to many textiles, even certain woolen ones, but, when large amounts are used, it is generally to cover defects in the material.

Experiment 38. Object: To determine the amount of dressing in a fabric.

If material contains a large amount of dressing, it will lose its stiffness when rubbed between the fingers. For this test, use some pieces of cotton and linen material that will do so and some that are not much changed by rubbing, in order to compare the results.

Procedure: Weigh the pieces of material and then boil them in a 3% solution of hydrochloric acid. Rinse them in water, each one separately, and test the water for starch with iodin. (See page 289.) If starch is present, in any of the pieces of material, boil those in which it is again. When they are all starch free, dry and weigh them. The difference in weight will show the amount of dressing that was present.

Printing, Dyes, and Dyeing

The difference between the dyeing and printing of fabrics.—By dyeing is meant *the saturation of the fibers of a textile with a coloring substance;* by printing, *the impression of a colored design upon a fabric.*

Printed materials can be usually distinguished from dyed by examining the wrong side of the fabric, if the pattern does not show through, it is printed.

Material may be dyed either before or after it is woven.

Dyes.—A few dyes are prepared from inorganic matter, notably, khaki, which is made by adding chrome alum to a solution of iron, but they are nearly all procured from organic matter. In the olden days, dyes were made principally from vegetable matter, and a few, as cochineal, from animal matter, but, at the present time, coal tar is their usual source. In fact, it has been estimated, that over fourteen thousand colors are secured directly or indirectly from coal tar.

Coal-tar dyes are often spoken of as the aniline dyes, because the first ones obtained from this source were made from aniline—a coal-tar derivative—now, however, a large number are obtained from carbolic acid, naphthalene, anthracene, and various other products of coal tar.

Coal-tar dyes vary greatly in their nature, some having acid and others basic properties as well as other different characteristics.

Mordants and lakes.—On account of the differences in the nature of dyes and those of textiles, all fabrics will not be equally well dyed with the same kind of dye; *e. g.*, the protein matter in silk and wool will unite with substances in some dyes that will find nothing to combine with in the cellulose fibers of textiles of vegetable origin. In such case what is known as a *mordant* is used. By a mordant is meant *any substance that will combine with some constituent of the textile or which can penetrate its fibers and which also has an affinity for the dyestuff so that it will unite with it and thus procure the saturation of the fibers with an insoluble colored substance.* This insoluble compound formed by the use of mordants is called a *lake.*

Permanence of dyes.—Even with the use of mordants, all dyes will not be equally permanent, and their permanence or *fastness* will vary in different materials.

The more common auxiliary causes of change of color in material are washing, the action of the acids of perspiration, and exposure to the sun's rays.

The following tests are ones that are frequently used to test the resistance of dyes to such things.

Experiment 39. Object: To test the color fastness of materials.

Test for washing fastness.—Soak and rub a small piece of the material in boiling soap solution. Do this several times. If the color does not run and if comparison with an unwashed piece of material shows that it has not faded, the color is *fast to washing*.

Resistance to perspiration.—Soak a piece of the material to be tested in 30% acetic acid that has been heated to about 99° F. and then dry it, without rinsing, between parchment paper. When dry compare it with a piece of the material that has not been so treated.

Resistance to light.—Place the pieces of material to be tested where they will be in the direct sunlight a considerable portion of the time for a month. Compare them daily with similar pieces that are not so exposed. Colors that will stand such exposure for a month are considered *fast;* those which have undergone appreciable though not very great change, are said to be *fairly fast;* those which show considerable change in two weeks are classed as *moderately fast*, and those which are much changed, as *fleeting*.

The comparative permanence of the fundamental colors in different materials is about as follows:

Brown lasts well in cotton materials, but is likely to fade in linen and woolens.

Dark blue is generally a fast color in all the textiles, but light blue generally fades quickly.

Black lasts well in wool and silk, but not in vegetable fibers.

Red is fairly permanent in woolen and silk materials, but only fairly so in others.

Gray is fairly permanent in all materials.

Pink fades soon, but usually does so uniformly and becomes a pretty shade.

Green cannot be depended upon except in high-priced woolens and silks.

Lavender is likely to be a fleeting color in all materials.

CHAPTER XVIII

THE CHEMICAL CONSTITUENTS OF THE HUMAN BODY AND OF FOOD

Classification, Nature, and Uses of the Substances Composing Animal Bodies and Plants—Tests for Proteins, Starches, Sugars, Salts—Origin of Food Material.

In order that the machinery of the human body (*i. e.*, the heart, lungs, muscles, etc.) may be kept at work, it must be provided with fuel and with matter containing the same elements as itself which its cells can utilize for their building and repair. The elements that the body needs for these purposes are contained in many compounds that are very common in nature, but almost the only combinations that the human body can utilize are those formed in certain plants and in some of the lower animals. Though the tissues of plants and of animals are so unlike in appearance, the compounds of which they are composed are so similar in their chemical composition that they, including those constituting the human body, are classified under the same headings.

Classification of the substances composing plants and animals.—The various substances that enter into the composition of both plants and animals are classified as *organic* and *inorganic*. The organic compounds, are further classified as *nitrogenous* (those containing nitrogen) and *non-nitrogenous* (those without nitrogen). The majority of nitrogenous substances are known as *proteins* and the non-nitrogenous organic substances are classed as *carbohydrates* and *fats*. The inorganic constituents of plant and animal tissues are salts and water.

Table showing classification:

Organic	Nitrogenous	Proteins and a few non-protein compounds that contain nitrogen
	Non-nitrogenous	Fats Carbohydrates
Inorganic		Salts Water

Examples of proteins are: The substances in egg which solidify when heated; the substance in milk that forms as a scum when the latter is heated and that which clots or curds when rennin or acid is added to milk; the white substance which appears as a coating on meat when it is exposed to heat.

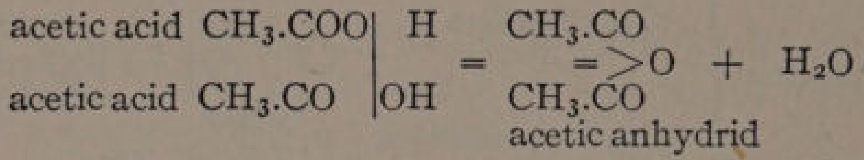
Familiar examples of carbohydrates are starches and sugars.

Proteins

The proteins are indispensable constituents of animal and plant cells; without them all life would cease.

Composition of proteins.—Protein molecules have a very complex structure, knowledge of which is still very limited, though it has been greatly added to lately. Proteins consist of the elements carbon, hydrogen, nitrogen, oxygen, sulphur, and sometimes phosphorus and iron. The more complex proteins contain thousands of atoms of these elements. The non-mineral elements of protein molecules are so combined that they form what are known as *anhydrids of amino acids*. These anhydrids are formed by reaction between the OH groups of two molecules of organic acid with a loss of water.

Example:



$$\begin{array}{l|l} \text{acetic acid } CH_3.COO & H \\ \text{acetic acid } CH_3.CO & OH \end{array} = \underbrace{\begin{array}{l} CH_3.CO \\ CH_3.CO \end{array} \!\!> O}_{\text{acetic anhydrid}} + H_2O$$

Amino acids.—These differ from other organic acids chemically in that their hydrogen is replaced by what is known as amine or the amino group or radical, which is NH_2. For example, acetic acid is represented by the formula CH_3 CO_2. H and amino acetic acid by the formula CH_3CO_2. NH_2. Amine is somewhat similar to ammonia—NH_3—and ammonia is formed from it during the decay of protein-containing compounds.

Classification of proteins.—A very large number of different amino acids have been obtained both by the decomposition of organic compounds and by synthesis, and as different kinds of proteins contain different numbers and kinds of these acids it can be appreciated that there is considerable variation in the nature of the protein of different foods.

Certain proteins, however, have similar characteristics and according to these they have been classified as follows:

- Simple proteins
  - Albumins
  - Globulins
  - Glutenins
  - Alcohol solubles
  - Albuminoids
  - Histones
  - Protamines
- Conjugated proteins
  - Nucleoproteins
  - Glycoproteins
  - Phosphoproteins
  - Hemoglobins
  - Lecithoproteins
- Derived proteins
  - Primary derivatives
    - Proteans
    - Metaproteins
    - Coagulated proteins
  - Secondary derivatives
    - Proteoses
    - Peptones
    - Peptids

Simple Proteins

Simple proteins are those which, when digested or otherwise hydrolyzed,[1] yield only amino acids or their derivatives.

Albumins and globulins.—Proteins belonging to these two classes are found in many of the same substances. As a rule, however, there are more albumins than globulins in animal fluids, as blood, and more globulins in animal tissues and plants.

The albumins and globulins have many characteristics in common. They are both colloidal and will not diffuse through animal membranes and, under the influence of heat, certain salts (as bichlorid of mercury, silver nitrate, etc.), and various other substances,

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[1] By hydrolysis is meant, *the splitting of complex molecules due to their absorption of water.*

their molecules tend to aggregate together to form a coagulum. They are both soluble in dilute salt solutions, dilute acids and alkalies, but albumins are also soluble in distilled water and concentrated sodium chlorid solutions and globulins are not. The classification of the simple proteins placed under these two headings was based largely upon these differences in their solubilities.

The several different kinds of albumins and globulins are given different names. Thus, albumin in serous fluids and in blood is called serum-albumin; that in eggs, ova-albumin; that in milk, lact-albumin; that in muscle tissue, myogen. The globulins of the blood are called fibrinogen and serum-globulin or paraglobulin; that in muscle tissue, myosin[1]; also there are a few albumins and various globulins in plants, but as their special names are not often used, they need not be given here.

The glutenins and alcohol-soluble proteins.—The glutenins and the most common alcohol-soluble protein, gliadin, form the gluten of wheat flour, which is its principal nitrogenous constituent. Gluten is found also in other cereals, but, as will be seen later, that in the other cereals is not quite the same as that of wheat flour.

Albuminoids.—These are protein substances extracted from such matter as bone, cartilage, and similar substances. They form the basis of gelatin.

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[1] Myosin, it is thought, does not exist as such in the living muscle, but is formed there by coagulation after death. The change is thought to be similar to the transformation that occurs in fibrinogen when blood clots. *Rigor mortis*—the stiffening of the body after death—is due to the changes that occur in the proteins after death.

The composition of the albuminoids differs from that of the other simple proteins, and gelatin, though it contains nitrogen, cannot be used in the diet as a substitute for them, since it cannot be utilized in the body, as other proteins are, for the building of muscle tissue.

Histones and protamines.—These two simple proteins are not common, they have been obtained from the blood corpuscles of some animals and from the spermatozoa of fish.

Conjugated Proteins

Conjugated proteins are those which have a molecule or molecules of matter other than mineral united to the protein molecule.

Nucleoproteins.—These are compounds of simple proteins and nuclein or nucleic acid. They are contained in the nuclei of cells and are relatively abundant in such glandular tissues as the spleen, pancreas, and liver, the cells of such organs containing many nuclei.

Nucleic acid, on oxidation in the body, yields substances known as the *purin bodies*, *e. g.*, uric acid. For this reason, foods containing a high per cent. of nucleoprotein are eliminated from the diet in diseases, such as gout, which are characterized by the presence in the system of a comparatively large amount of uric acid.

Glycoproteins.—These are proteins which contain a carbohydrate molecule attached to the protein. Mucin, which is a constituent of the secretions of certain glands and of the mucous membranes of the respiratory and alimentary tracts, is an example.

Phosphoproteins.—These are compounds of the protein molecule and some phosphorus-containing substance or substances. Caseinogen of milk (the protein that is clotted when rennin or acid is added to milk) and vitellin, the principal protein of egg-yolk, are two of the most important proteins of this class.

Hemoglobin.—This is a compound of a protein called *globin* and a pigment (coloring matter) called *hematin* that unites and parts with oxygen readily. Hemoglobin is the essential constituent of the red blood-corpuscles of the blood, for it is due to their hemoglobin that the corpuscles can fulfill their function of carrying oxygen from the lungs to the tissues. Hemoglobin contains iron and this constituent is necessary both for its formation and functioning.

Lecithoproteins.—These are compounds of protein and lecithin, a peculiar phosphorized fat. They are contained in nerve tissue, in mucous membranes, and probably in small amounts in other tissues. Yolk of egg is the food material in which they are found in largest amounts.

Derived Proteins

Such proteins are derived from simple and conjugated proteins as the result of digestion and they may be also obtained by boiling simple or conjugated proteins in acid.

The primary derivatives are those which have not undergone any great change from their original form. Secondary derivatives are those in which the change has been carried to a greater degree. They are not coagulated by heat.

Function of food proteins.—Protein food substances are oxidized in the body, consequently they yield heat and energy, but their special function is to provide building and repair material for the proteins of the body.

Non-protein nitrogenous compounds.—There are certain nitrogen-containing compounds in both plant and animal foods that are not proteins. They have little or no nutritive value, but they give flavor to the food and stimulate the secretion of the digestive juices. Examples of these compounds are the amids of plants and the extractives of meat.

The amids are believed to be the form in which nitrogen compounds are transferred from one part of the plant to another. The extractives of meat are substances formed in the animal body in the course of metabolism.

Tests for Proteins

Certain tests are often used to discover if protein substances are present in compounds. Some of these tests consist in adding something to the matter to be tested that will, if protein be present, unite with it and form a colored compound. Others consist in the addition of a substance that will coagulate the protein if it be present, or in the application of heat, which will have a similar effect. As will be seen in the experiments connected with the study of digestion, derived proteins do not respond in the same manner as the undigested proteins to all of these tests.

Three common tests which depend upon the formation of colored substances are the Piotrowski, Rose, or biuret; the xanthoproteic; and the Millon.

Methods of Performing Tests

Biuret. To 3 c. c. of the solution to be tested (in a test tube) add an equal amount of sodium hydroxid and a drop of a 1% solution of copper sulphate. If protein is present, a violet or pink color is produced, which, depending upon the nature of the protein. See page 398.

Xanthoproteic. To 5 c. c. of the solution to be tested add about one-third of its volume of strong nitric acid. If protein is present, a white precipitate will form. Boil the solution for one minute; the precipitate will turn yellow and partly dissolve; cool and add enough ammonia to make the solution alkaline; an orange color indicates positively the presence of simple or conjugated proteins.

Millon. Treat 5 c. c. of the solution to be tested with half its volume of Millon's reagent (see page 20). A white precipitate is formed. Boil the mixture. If protein is present, the precipitate turns brick red in color or disappears and leaves a red solution.

Coagulation of Proteins

The majority of proteins are coagulated by heat; important exceptions are albuminoids, the casein of milk, proteoses, and peptones. Heat is therefore sometimes used as a test for the presence of proteins in liquids, especially urine. When the heat test is used, a drop or two of dilute acetic acid is usually added after the liquid has been heated, because it intensifies the coagulation and also dissolves earthy phosphates, which may also be present, and which are likewise precipitated by heat.

Dried protein matter is not coagulated as easily as that containing water, and this fact has two important bearings on the destruction of bacteria, since heat kills bacteria by coagulating their protein constituents: (1) Dry heat causes evaporation of the moisture from bacteria and the soft matter surrounding them, which is one reason why it does not penetrate as thoroughly nor act as quickly nor at as low a temperature as moist heat. (2) Spores contain much less water than the rest of the bacterial protoplasm and are thus much more resistant to heat.

Heat coagulation will be discussed further in connection with the chemistry of cooking.

Proteins are coagulated also by: (1) the salts of heavy metals, as bichlorid of mercury, nitrate of silver, copper sulphate, lead acetate; (2) strong solutions of alcohol (solutions of alcohol between 10 and 50 per cent. will precipitate proteins, but they will not, as do stronger solutions, form an insoluble coagulum); (3) alkaloidal reagents, as potassium ferrocyanid; (4) tannic acid and inorganic acids.

It is due to their effect on proteins that bichlorid, silver nitrate, etc., act as poisons and as disinfectants, since coagulation of the protein of living organisms, be it of bacteria or human beings, will cause death. Bichlorid cannot be used for the disinfection of substances containing much protein matter, because the coagulum that is formed with protein is not a disinfectant and it forms such a hard case around the bacteria that they are protected from the influence of the mercury. Alcohol 95 per cent. has no germicidal action but a 70 per cent. solution will destroy vegetative forms of bacteria in from fifteen to ten minutes or less. The failure of the 95 per cent. to act as a disinfectant is due to the formation of a protecting envelope, by the coagulation of the outer protein substance of the germs through which the alcohol cannot pass to destroy their vital principles.

The difference of the nature of the coagulum produced by different substances can be studied by combining the white of egg and about 300 c. c. water, putting about 5 c. c. of this solution into several test tubes, adding a few drops of a different coagulating reagent to each tube and carefully examining the comparative toughness and testing the solubility of the precipitates in water and dilute acid solutions.

Though precipitation is often spoken of as coagulation, a precipitate that is readily dissolved is not a true coagulum.

The clotting of milk and blood.—When milk is treated with rennin and kept at body temperature, a clot or curd is formed, due to the conversion of the soluble caseinogen of the milk into an insoluble protein called *casein*. Likewise when blood is shed and, under some abnormal conditions, in the living blood-vessels, the soluble blood protein, fibrinogen, is changed to an insoluble protein called *fibrin* and forms a clot. This change in the fibrinogen is due to the action of a ferment called *thrombin*, which is formed from a substance called *thrombokinase* (that is liberated from the blood-platelets and leucocytes on their disintegration) by the influence of the calcium salts of the blood.

The clotting of milk will be further referred to in connection with the chemistry of cooking.

The clotting of milk and blood is often spoken of as *coagulation*, but some authorities object to the latter term being used for this reaction since it produces entirely new protein substances.

The Carbohydrates

Classification.—The carbohydrates are classified as follows:

Monosaccharids	Glucose or dextrose
	Fructose or levulose
	Galactose
Disaccharids	Sucroses
	Lactose
	Maltose
Polysaccharids	Celluloses
	Starches
	Dextrins
	Glycogen
	Gums
	Pectins

Constituents.—The carbohydrates are composed of carbon, hydrogen, and oxygen. The proportions in which these elements exist in the molecules of the different carbohydrates are as follows:

Monosaccharids—$C_6H_{12}O_6$
Disaccharids—$C_{12}H_{22}O_{11}$
Polysaccharids—$C_6H_{10}O_{5x}$

Polysaccharids—Gr. *poly=many;* Lat. *saccharum =sugar.*—The carbohydrates of this group are considered chemically as being complexes of the simple sugars or monosaccharids, for though, up to this time, the chemist has not succeeded in building cellulose, starch, etc., from sugar, these polysaccharids are made from sugar in plants, and starches are easily changed to sugar in the course of digestion and in other ways that will be discussed later.

Cellulose

Cellulose is, next to water, the most abundant constituent of plants for it forms much of their framework. The nature of cellulose varies in different plants and in most plants at different stages of growth. Thus the cellulose walls that hold the starch grains of the potato are less fibrous than the stem of the celery plant, and that of the latter is less tough than the cellulose of the wood of trees. The cellulose of young plants is more tender than that of older ones, the cell walls of most plants becoming thicker and tougher as they grow, due to increase in the deposition of cellulose and decrease in the amount of water.

Celluloses are insoluble in water, either hot or cold, and in weak acids and alkalies, but strong acids and alkalies hydrolyze them. When celluloses are partially hydrolyzed, they react to the iodin test in the same manner as the starches do. Nitric acid converts celluloses into nitro-celluloses, which are the substances used as bases in the preparation of such matter as explosives, collodion, and nitro-glycerin.

Digestibility of cellulose.—Some of the lower animals apparently digest cellulose quite readily, but, though the more tender forms are to a certain extent permeable to the digestive fluids of the human alimentary canal, all cellulose is very insoluble, and in man whatever digestion of this substance does take place occurs in the intestine, largely as the result of bacterial action, and fatty acids seem to be the principal product yielded. Though a large portion of the cellulose of food may thus escape digestion, a certain amount of it is a valuble adjunct to the diet of normal individuals, except very young children, because it stimulates peristaltic action in the intestines and thus prevents constipation.

Starch

Starch is the principal form in which plants store their food supply. It is stored more especially in the roots, tubers, fruit, and seeds.

Properties of starch.—Starch is insoluble in water. When heated, starch grains swell and rupture and prolonged exposure to heat causes chemical changes which render them soluble. The substance thus produced is known as *dextrin*. If heated with acid and as the result of digestion, starch is changed to dextrose.

When dry, starch will keep in good condition for a long time, but it is readily attacked by micro-organisms if it becomes moist, and in such case will mould or ferment and become sour.

Starch is obtained chiefly from wheat, corn, rice, and potatoes; and arrowroot, sago, and tapioca are nearly pure starch.

Test for Starch

A test very commonly used to detect the presence of starch in a substance is the iodin test. It is as follows: A drop of iodin solution, 3%, is added to the solution to be tested. If the material to be tested is solid, it is boiled for a minute with water, so that it will be in solution, and cooled before the iodin is added. If starch is present, a blue color will develop, due to the formation of iodid of starch. (N. B.—The reaction will not take place properly if the material being tested is hot.)

Dextrin

Occurrence.—Dextrin is only occasionally found in uncooked foods. It is formed when starch is heated to about 200° C. (the crust of bread is an example) and when starch is boiled. It is the first product formed in the digestion of starch.

Test for Dextrin

The iodin test, as used for starch, is also frequently used to detect the presence of dextrin in a substance. If dextrin is present in the material tested, the iodin will turn either purple, red, or white. The reason for this difference in color is given on page 398.

Glycogen

Nature and occurrence.—Glycogen, known also as animal starch, is a white powder that is soluble in water. It is extracted from liver. It is formed in the animal body, in the liver, from glucose, and is stored in the liver and, to some extent, in the muscles, but that in the muscles is decomposed soon after death; therefore, liver is the only meat that contains any amount of carbohydrate. Oysters and other shellfish, however, contain as much as 9 per cent. glycogen.

Inulin

This is a soluble starch-like substance that is found dissolved in the sap of some plants, especially the Jerusalem artichoke.

Pectin

Pectin substances are related chemically to both starches and gums. It is thought that they are formed in plants by the combining of several simple carbohydrates. It is the pectin substances that, under conditions which will be discussed later (page 362), allow of jellies being made from certain fruits without the aid of gelatin. The test for pectin will be found on page 362.

Gums

These are various viscid substances contained in many plants and in the wood and bark of several trees. They have the same chemical composition as the other polysaccharids.

Disaccharids

The sugars belonging to this class are the sucroses, lactose, and maltose.

The disaccharids were so called from the Gr. *di* = *double* and Lat. *saccharum* = *sugar*, because, when hydrolyzed, one molecule of a disaccharid gives two molecules of simple sugars or monosaccharids.

Sucroses

The more common sucroses are cane sugar, beet sugar, and maple sugar.

Sucrose is a common constituent of plants. It occurs in small quantities in several fruits, usually in combination with dextrose and levulose, and in a few vegetables. It is found in largest amounts in the sugar cane, beets, the maple tree, pineapple, and carrots. It is from the three plants mentioned first that it is usually extracted.

Sucrose is easily hydrolyzed in digestion and by boiling in water or other liquids, more especially acid solutions. When hydrolyzed every molecule of sucrose yields a molecule of dextrose and a molecule of levulose. Sucrose can be changed to caramel by the application of dry heat.

Test for Sucrose

The presence of sucrose in food can be determined by boiling a few c. c. of a solution of the matter to be tested with an equal quantity of hydrochloric acid. If sucrose is present, a deep red color will develop.

Lactose

Lactose, known also as sugar of milk, occurs in the milk of all mammals. Experiments seem to show that it is made in the mammary glands from glucose which secretory cells of the glands take from the blood.

Lactose is easily changed to the monosaccharids glucose and galactose by heating it with acids, and in digestion. Lactose will reduce Fehling's solution, but it requires 0.68 gm. of lactose to reduce 10 c. c. of Fehling's. As will be seen later (page 398), one method of determining the nature of a sugar is to ascertain if it will reduce certain solutions and how much of the sugar is required to cause the reduction of a given amount of the solutions.

Maltose

Maltose or malt sugar is formed from starch by the action of diastatic enzymes (see page 389) that exist in certain plants. It is therefore an ingredient of germinating cereals, of malt and malt products, and in animals it is formed during digestion. It is easily changed to glucose by heating it with acids, by the action of yeast, and in digestion.

Maltose is the principal constituent that the brewer extracts from malted grains for the making of beer, ale, etc. In the process to which it is subjected it is changed to glucose, and this is easily oxidized, thereby forming alcohol and, if the process is continued, CO_2 and H_2O.

Monosaccharids

The monosaccharids were so named from the Gr. *monos* = *one* and Lat. *saccharum* = *sugar*.

Dextrose

Dextrose was so named from the Lat. *dexter* = *right*, because it turns polarized light to the right. It is known also as *glucose*, *diabetic* and *grape sugar*. It is found in combination with other sugars, especially levulose, in many fruits and a few other plants.

The disaccharids and starches are easily changed to glucose by boiling them with acids and much of the glucose sold is made from starch in this way. Glucose is only about half as sweet as cane sugar.

Normally, glucose is present in the blood in the proportion of 0.1 per cent. This will be further discussed in the sections on digestion and metabolism.

Levulose

Levulose (Lat. *lævus=left*) so called because it turns polarized light to the left, is known also as fructose and fruit sugar. It is an isomer of glucose—*i. e.*, it contains the same elements in the same proportions as glucose, but in a different combination. The way in which the chemist demonstrates this difference of arrangement is shown on page 157. Levulose is not found in the blood, other than that of the portal vein, being changed into glycogen in the liver. It is found associated with glucose in fruit and in honey.

Galactose

This monosaccharid is not found in nature, it is formed from lactose by hydrolysis.

Tests for Monosaccharids

It is often important to discover if glucose is present in a food or other matter—*e.g.*, the urine. The Fehling's test is one very commonly used for this purpose.

Fehling's test.—Take equal parts of Fehling's solutions 1 and 2 (see page 19), about 2 c. c. of each. Boil[1] this, if it remains clear, add an equal amount of the solution to be tested, and boil. If glucose is present, a red precipitate will form; because the glucose by taking away oxygen from the cupric hydroxid changes it to the red insoluble cuprous oxid.

Other common tests are as follows:

Barfoed's test.—Boil a few c. c. of freshly prepared Barfoed's reagent and add, while the solution boils, drop by drop, some of the solution to be tested. If glucose is present, a red precipitate of cuprous oxid is formed, either at once or on standing.

Benedict's test.—To 5 c. c. of Benedict's reagent, in a test tube, add about 8 drops of the solution to be tested. Boil the solution for about 2 minutes and then allow it to cool. If glucose is present, the solution will become filled with a precipitate either red, yellow, or green in color, depending upon the concentration of the sugar.

These tests all depend upon the fact that the monosaccharids have the same chemical construction as the aldehydes and like them will reduce such substances as the reagents used in these tests, thereby changing certain soluble substances into insoluble ones with the colors mentioned.

Moore's test.—Boil the solution to be tested with a little sodium hydroxid. If glucose is present, the solution will turn yellow, due to the formation of caramel as the result of the hot alkaline on the glucose.

Fermentation test.—Fill the long arm of a fermenta-

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[1] This precaution is necessary because the solution often deteriorates soon after it is prepared. If, when boiled, it becomes cloudy, it is not fit to use.

tion tube (page 19) with the solution to be tested and add a small piece of yeast cake. The liberation of gas (carbon dioxid), as described on page 359, indicates the presence of sugar.

Fats and Oils

As stated in Chapter XIV., fats are esters of glycerin and fatty acids.

They are classified as:

Fats
Fixed oils
Essential or volatile oils.

The essential or volatile oils.—These oils are so called because they are volatilized at ordinary temperatures. If a little volatile oil and a little fixed oil are dropped on paper, the latter will leave a mark, but the former, being volatile, will quickly disappear.

The essential oils vary considerably in their composition, but the majority of them consist chiefly of such substances as the terpenes, or camphors, or similar substances. They are all entirely different in their chemical composition from the fixed oils.

The essential oils are the important constituents of spices and similar substances; it being to their oils that these things owe their flavor. As the oils are volatile, they can be extracted from the plants containing them by distillation. Oil of cloves, oil of peppermint, and other oils used for flavoring are obtained in this way.

Nature of the fixed oils and fats.—The fixed oils and fats are chemically alike and their classification is based on their physical condition: fats being solid at ordinary temperatures and oils liquid, but a fat when melted is called an oil, and an oil when solidified a fat.

Fixed oils and fats are derived from both plants and animals. They are composed of three or more simple fats. The molecules of the simple fats, it will be remembered, consist of three molecules of fatty acid and one molecule of glycerin.

The most common simple fats are stearin, palmitin, and olein. The first two mentioned are solid at ordinary temperatures, and olein is liquid, therefore the greater the amount of olein in a fat, the lower its melting point. The fats of beef and mutton, for example, contain a larger per cent. of stearin than lard and butter and are therefore harder.

Nearly all animal fats are mixtures of these three fats, but the fat of milk (cream) and consequently butter contains several other simple fats of which the principal one is butyrin. Plant oils contain little or no stearin, but many of them have several other simple fats.

Solubility.—Fixed oils and fats are insoluble in water, but they are readily dissolved in hydrocarbons, ether, carbon bisulphid, chloroform, and hot alcohol. They are saponified by hot alkalies and in digestion.

Function.—The function of fats in the body is to provide heat and energy and material for building fatty tissue.

Lecithins

These are substances related chemically to the fats. They are compounds of glycerophosphoric acid radicals and a nitrogenous base called *cholin*.

Lecithins appear to be an active component of both animal and vegetable cells, but they are especially abundant in the white matter of nervous tissue and in the yolk of eggs.

Lecithins are digested in the same manner as the true fats. Their function is not understood, but it is thought that they may assist in the synthesis of the nucleoproteins.

Mineral Matter

The mineral matter of food and the human body is spoken of also as *ash* and *salts*. It is called *ash* because, being non-combustible, it is found in the residue remaining when animal and vegetable substances are burned. The compounds found in the ash are not necessarily exactly the same as those contained in plant and animal tissues, for during ignition many chemical changes and consequent transformations occur. The term salts is used because the mineral matter occurs in plant and animal tissues chiefly in the form of salts.

The salts present in largest amounts in food and the human body are the chlorids, sulphates, phosphates, and carbonates of potassium, sodium, calcium, and magnesium.

Function of salts in the body.—As salts are not oxidized in the body they do not yield heat and energy, nevertheless they are necessary for life; for, they are essential constituents of protoplasm; they maintain the neutrality or slight alkalescence of the body secretions, blood, and other fluids; they are necessary for the maintenance of the elasticity and irritability of muscle tissue; calcium salts are the source of the firmness and rigidity of the bones;

sodium chlorid is the substance from which the hydrochloric acid of the gastric juice is made; iron is essential for the making of the hemoglobin of the red blood-corpuscles and for the oxygen-carrying power of the latter and it is an equally important constituent of the chlorophyl of plants.

Solubility of salts.—All the chlorids and the potassium and sodium sulphates and phosphates are soluble in water; the normal phosphates of calcium and magnesium are insoluble in water, but soluble in dilute acids.

Common Tests for Determining the Presence of Salts in Food and Other Matter

To about 5 c. c. of the solution to be tested add a few drops of strong nitric acid and then a few drops of a 1% silver nitrate solution ($AgNO_3$). If chlorids are present, a white flocculent precipitate will appear; because, in the presence of nitric acid, the chlorin of the chlorids unites with the silver, forming silver chlorid which is insoluble in water and dilute acid solutions. Write the equation for the chemical reaction that occurs. (The nitric acid does not enter into the reaction.)

Test for iron.—To about 5 c. c. of the solution to be tested, add a drop of nitric acid and then 2 c. c. of potassium thiocyanate (KSCN). If iron is present, a deep red color will develop.

Test for phosphates.—To about 5 c. c. of the solution to be tested, add a few drops of (1) concentrated nitric acid and (2) ammonium molybdate solution, $(NH_4)_2MoO_4$. If phosphates are present, a yellow precipitate of ammonium phospho-molybdate is formed.

Test for sulphates.—To about 5 c. c. of the solution to be tested, add a few drops of HCl and then, slowly, about 2 c. c. of a 5 per cent. solution of barium chlorid ($BaCl_2$). If sulphates are present, a white insoluble precipitate of barium sulphate is formed.

Before testing a solid substance for mineral matter, it must be incinerated (*i. e.*, burnt until it is reduced to ashes[1]) and the ash dissolved in water, or, for the extraction of salts that are insoluble in water (see page 299), water plus a little concentrated nitric acid solution. This is filtered and it is the filtrate that is tested, since the salts, being soluble in the liquid for their extraction, will be present in it in solution.

Water

Functions.—Water, like the mineral matter of food, has no fuel value, but it is an essential constituent of all protoplasm, and it enters into the composition, not only of the fluids of animals and plants, but even of such hard substances as bone and wood. The water of the blood of animals and the sap of plants is the vehicle for distributing food material through the organism and for carrying away waste matter to parts from which it can be eliminated. Water dissolves food so that it can be used by the body, and it is quite as essential to life as any of the other food materials. The plant bereft of water quickly perishes, and, though animal life will not succumb quite so quickly, water is as essential for the well-being of animal cells as of those of the plants; therefore water lost from the body, as by excessive perspiration, diarrhea, or hemorrhage, should be replaced as soon as possible.

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[1] The usual method of doing this in a laboratory is to place the material in a crucible and put this over a flame.

Origin of Food Material

Photosynthesis (Gr. *photos = light* and *synthesis = composition*).—Photosynthesis is the name given to the processes whereby, under the influence of energy derived from the sun and chlorophyl (the green coloring matter of plants), cells in the leaves of green plants put together CO_2, which they obtain from the air, and H_2O, which they absorb from the soil, to form sugar and starches.

The exact nature of the reactions that occur in these processes is not known, but the following equations show how starch could be made from these compounds:

$$12CO_2 + 12H_2O = 2C_6H_{10}O_5 + 24O + 2H_2O$$

The manufacture of starch, however, is probably seldom, if ever, such a simple matter as this. It is thought that sugars of the monosaccharid type are formed first and the more complex sugars and starches made from these. This could happen as follows:

$$6CO_2 + 6H_2O = C_6H_{12}O_6 + O_{12}$$
$$C_6H_{12}O_6 - H_2O = C_6H_{10}O_5$$

but, even in the manufacture of simple sugars, it is considered that certain preliminary steps, in which formaldehyde and peroxid of hydrogen are formed, may be usual. Thus

$$CO_2 + 3H_2O = \overset{\text{Formal-dehyde}}{CH_2O} + \overset{\text{Hydrogen peroxid}}{2H_2O_2}$$
$$6CH_2O = C_6H_{12}O_6$$

Metastasis (Gr. *meta = beyond* and *histanai = to place*).—In addition to sugars and starches, plants must make cellulose (which constitutes the basis of their own structures), fats, and proteins. Cellulose is chemically very similar to starch. Fats and proteins have the same elements as starch, but in very different proportions, and proteins contain in addition N, S, and sometimes P and Fe. These elements plants obtain from the earth in salts, such as nitrates, sulphates, etc. Still less is known of the processes that occur in the making of fats and proteins than of those by which sugar and starch are formed, but it is thought that these substances are made from starch and that the latter is broken down into simpler substances before synthesis occurs. The forming of proteins and other complex compounds, including the destructive processes necessary for their formation, are spoken of as *metastasis*.

Photosynthesis takes place only in the green leaves of plants and in the daylight or under strong electric lights, but metastasis takes place in all parts of growing plants and in the darkness as well as in the light.

It is through their leaves that plants absorb the CO_2 which they require for their food and a small amount of water enters in this way, but by far the greatest amount of their water supply is obtained from the soil, being absorbed by the surface cells of their roots and root-hairs. This water has in solution salts, including the nitrates, etc., required for the manufacture of proteins.

The nature and purpose of sap.—The sap of plants consists of water holding in solution the salts absorbed from the earth and the food material that was made in the principal food factory—the leaves—some of which must be carried to the storehouses—the roots, seeds, fruits, or flowers—for storage and some to all the cells of the plants for their nutrition. Thus, sap is to the plant what blood is to the animal, and, like blood, it must circulate throughout the organism.

The forces which cause the circulation of sap.—The factors which cause the movement of sap in plants are still but imperfectly understood, but five important forces are: (1) the avidity that protoplasm has for water; (2) the life processes of the plant; (3) osmosis; (4) evaporation; (5) capillarity.

Due to the avidity that protoplasm has for water, it will absorb water until it is saturated; therefore, as the water in the cells of the leaves evaporates or is used up in the life processes of the plant—*e. g.*, the manufacture of sugar—the uppermost cells absorb water from those beneath them, and these, in turn, take it from those which they overlie, and so on. This constant evaporation and absorption of water acts very much as a suction pump and it assists the factors upon which osmosis depends, see page 67, in causing the fluid in the soil to osmose through the root-hairs and roots and pass up the stem to the leaves, flowers, etc.

In warm, bright weather, evaporation and the chemical processes going on in the leaves occur much more rapidly than in cold weather, consequently the sap rises more rapidly and to a greater extent under the former conditions.

The sap must descend, as well as ascend, for the stems and roots need to be nourished and, as already stated, most of the nourishment is formed in the leaves. Less is known, however, of the forces which cause the sap to descend than of those which produce its ascent, but it is known that it descends through the large outer cells that, in the tree, constitute the bark. This is why removing its bark will kill a tree.

Source of salts in the soil.—It was stated in a preceding paragraph that plants got their nitrogen and mineral supply from the soil. This being the case, the soil must be constantly losing its supply of such matter and it must be as constantly renewed if plants are to grow in it. This is done by the use of fertilizers. Many compounds of the salts necessary for fertilizing the soil are now prepared by the chemist, but, until recently, animal excreta and decaying vegetable matter were the only fertilizers to be had and they are still very much used.

To understand how animal excreta can supply the soil with mineral matter it is necessary to realize that animal excreta contain not only salts eaten in food, but also, especially the urine, salts formed in the body by chemical reactions occurring in metabolism.

The protein in plants and in animal excreta does not exist in the form of nitrates, but when decaying vegetable matter and manure are spread over the ground they, by the influence of bacteria, undergo decomposition, and their protein matter is slowly formed into ammonia compounds, some of which, uniting with oxygen, and in other ways, are broken down to nitrates and other nitrogen salts which are capable of going into solution in the soil water and being absorbed by plants.

Another way in which bacteria supply plants and the soil with nitrogen is that various species are able to absorb nitrogen from the air and some of these species enter the roots of such plants as peas, beans, clover, alfalfa, and they live and multiply in little tubercles that develop on the roots in consequence of the invasion of the plants by these organisms. The bacteria, as they develop in these nodules, in some unknown way, combine the nitrogen they absorb so that it is available for plants. It is thought that perhaps the bacteria after absorbing the nitrogen secrete a nitrogenous substance, some of which is absorbed by the plant cells while a portion osmoses into the soil. The nitrogen supply of the soil in which such plants grow is thus increased, instead of diminished, as it is when used for the growth of other plants. The soil nitrogen supply is still further increased when the roots of such plants are allowed to rot in the ground.

The food cycle.—As shown in the preceding pages, plants and animals, including man, are absolutely dependent upon each other for their sustenance, and there is a continual cycle in progress by which one form of life supplies the other with matter for their life and growth. Animals eat plants and, when these are oxidized in their bodies, they eliminate CO_2 and other waste matter that plants can use for their food and development. Man supplies CO_2 also by burning plants, as wood, and plant products, as coal, oil, and gas.

Effect of plant life upon the atmosphere.—Plants not only supply man and other animals with food, but they also keep the air pure by extracting CO_2 and supplying oxygen. One source of the oxygen supply can be seen in the equations on page 301. Oxygen is freed also in the making of fats for, as can be seen in the following formulas of some of the simple fats, there is less oxygen in fat than in sugar and starch:

Stearin—$C_3H_5(C_{18}H_{35}O_2)_3$
Olein—$C_3H_5(C_{18}H_{33}O_2)_3$
Palmitin—$C_3H_5(C_{16}H_{31}O_2)_3$

In some metastasis processes CO_2 is liberated; therefore plants eliminate a small amount of CO_2, especially at night when photosynthesis is not taking place.

CHAPTER XIX

NATURE AND NUTRITIVE VALUE OF SOME OF THE COMMON FOODS

Classification of Foods—Nature, Digestibility, and Nutritive Value of the more Common Foods and Beverages—Nature and Action of Condiments and Spices.

Animal Foods

THE more common animal foods include the flesh of mammalia—generally spoken of as *meat*—birds, fish and other sea-food, eggs, milk, and honey.

Nutritive value of animal foods.—With the exception of honey, milk, liver, and shellfish, there is little or no carbohydrate found in animal foods. The reason for the absence of carbohydrate in meat is that that which is present in the living animal is drained off with the blood when the animal is slaughtered, or else changed to lactic acid and other products of the decomposition of glycogen and glucose shortly after the death of the animal. Consequently the only solids provided by the majority of animal foods are proteins, fats, and mineral matter.

Animal foods are usually a more expensive form of nutriment than vegetable foods, but, as can be seen by looking at the table on pages 340–344 they contain more protein than other foods, and animal protein is more thoroughly digested and absorbed than vegetable protein. Also, the flavor of meat, the stimulating effect of the meat extractives on the glands which secrete the gastric juice, the many uses of eggs and milk, all make animal foods very desirable articles of diet for the majority of people.

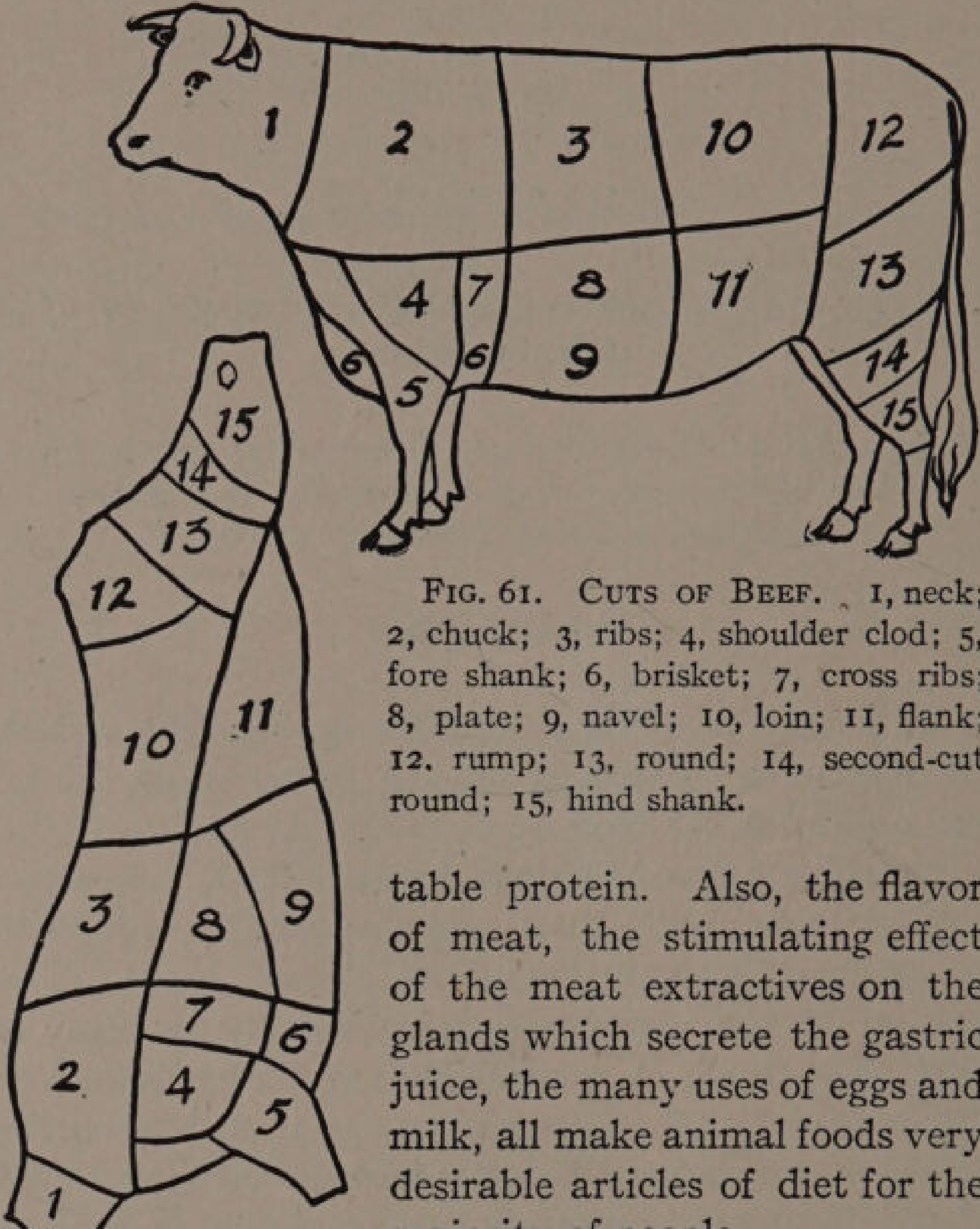


FIG. 61. CUTS OF BEEF. 1, neck; 2, chuck; 3, ribs; 4, shoulder clod; 5, fore shank; 6, brisket; 7, cross ribs; 8, plate; 9, navel; 10, loin; 11, flank; 12, rump; 13, round; 14, second-cut round; 15, hind shank.

Comparative digestibility of animal foods.—The following list of the comparative digestibility of common animal foods is one to be found in several books on dietetics. The foods are mentioned in the order of their digestibility, beginning with the most digestible.

Milk, oysters, soft-cooked eggs, sweetbreads, white fish, chicken, lean beef, scrambled eggs, mutton, squab, crisp bacon, fowl, tripe, lamb, corned beef, veal, ham, ducks and game, salmon, mackerel, herring, roast goose, lobster and crabs, pork; smoked, dried, or pickled fish and meats.

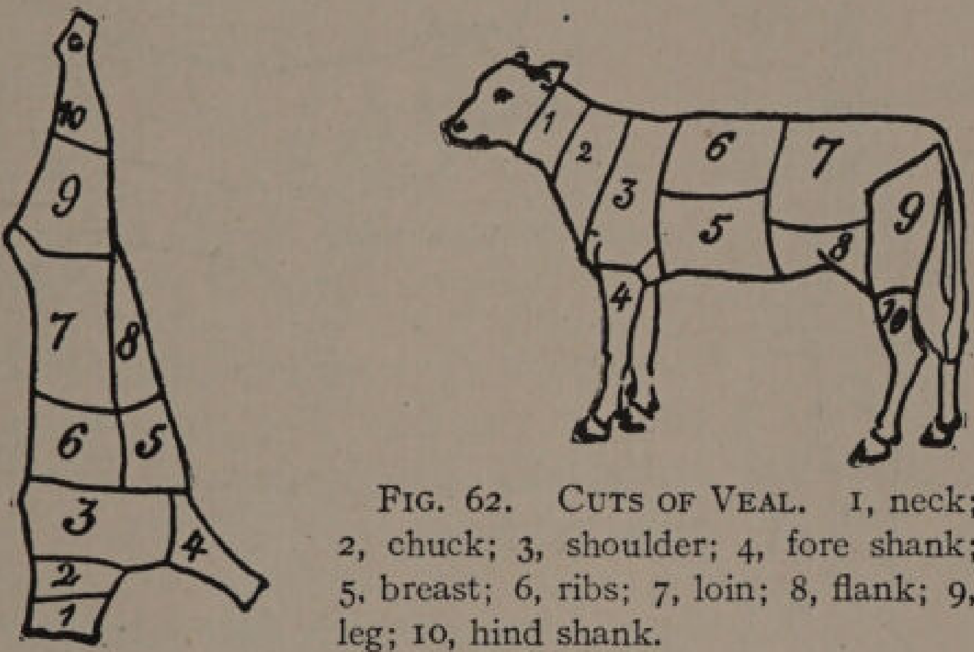
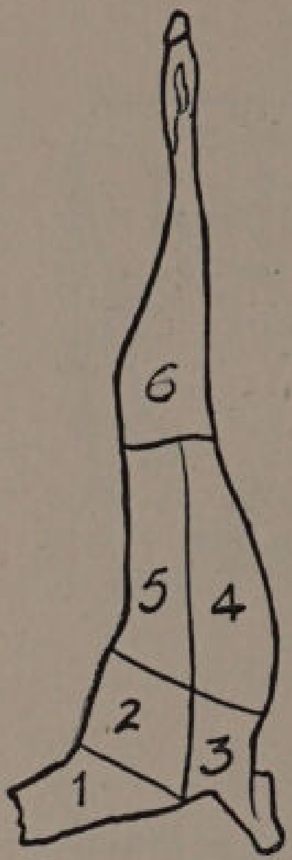


FIG. 62. CUTS OF VEAL. 1, neck; 2, chuck; 3, shoulder; 4, fore shank; 5, breast; 6, ribs; 7, loin; 8, flank; 9, leg; 10, hind shank.

Meat

Factors which influence the digestibility of meats.—In general the amount of fat and connective tissue and the length and strength of the muscle fibers which make up the muscular tissue (*i. e.*, the lean part of meat) are the determining factors of the digestibility of meat. If there is a large amount of connective tissue or if the fibers are long and coarse, the meat will be tough and, unless the condition is ameliorated in the cooking, difficult to masticate and otherwise digest. The fibers in certain cuts of meat are coarser than those in others, but the consequent toughness can be often rectified by proper cooking, and as a rule these cuts contain more nutriment and extractives than the more tender ones.

The fibers of the flesh of young animals are shorter and, consequently, the meat is more tender than that of older animals, but an excess of fat in lamb and a lack of extractives in veal make these meats, especially veal, less easy to digest than beef and mutton. The large amount of fat in pork interferes with its digestion. Bacon, however, is usually comparatively easily digested, partly as the result of the action of the salt used for its preservation upon the fat. With the exception of bacon, salted, pickled, and dried meats are not as easily digested as fresh meat for the fibers are hardened by the preservatives and by drying.



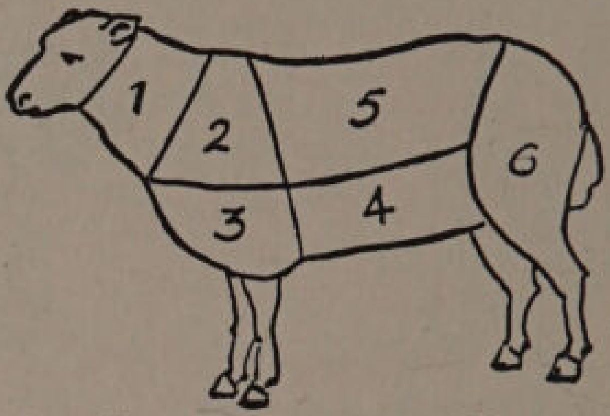
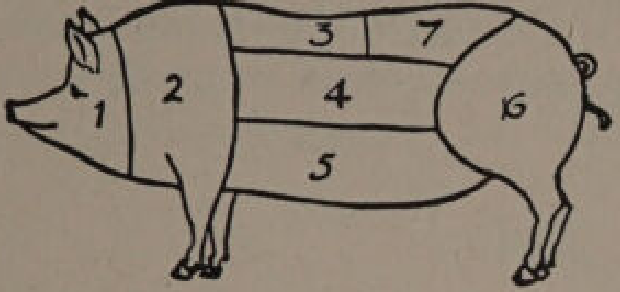


Fig. 63. Cuts of Lamb and Mutton. 1, Neck; 2, chuck; 3, shoulder; 4, flank; 5, loin; 6, leg.

Structure of meat.—Meat is made up of muscle fibers held together with connective tissue. Within the fibers are the so-called *meat juices* which consist of solutions of proteins, nitrogenous extractives, and salts in water. The extractives, which consist chiefly of substances such as creatin, xanthin, proteoses, and peptones,[1] give meat its flavor and they stimulate the secretion of the digestive juices, but the greater part of these substances are absorbed from the intestine and excreted by the kidneys without undergoing oxidation and thus they have no real nutritive value. Nevertheless, the flavor of meat depending upon these extractives, their preservation is one of the important considerations in the cooking of meat.



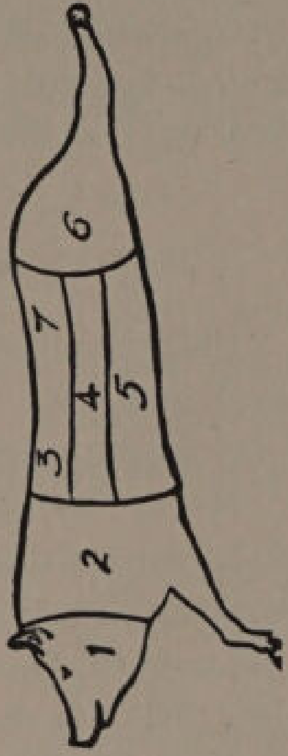


Fig. 64. Cuts of Pork. 1, Head; 2, shoulder; 3, ribs and back; 4, middle cut; 5, belly; 6, ham; 7, loin.

Birds

The flesh of birds is similar to that of meat, but the muscle fibers are shorter and therefore, especially in young birds, the flesh is more tender. Chicken, turkey, and squab are all easily digested, but ducks and geese have too much fat, and the majority of

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[1] Proteoses and peptones are not present in living muscle, but enzymes in the tissues, in the presence of the lactic acid, which accumulates in the body after death, cause the hydrolysis of some of the simple proteins, thereby giving rise to these derived proteins.

birds classed with game are too rich in extractives to be a suitable food for all invalids who are able to have chicken and squab. The presence of extractives, within limits, increases the digestibility of a food because they stimulate the secretion of digestive juices, but large amounts are to be avoided when there is any gastric disturbance.

Fish

Fish are usually classed as *scaly* or *vertebrate fish* and *shellfish*. The latter class includes the mollusks and crustaceans.

Nature and nutritive value of the scaly fish.—These fish contain more water and rather more mineral matter, especially the phosphates of calcium and potassium, than meat, also they contain more gelatin, but less fat and extractives. Consequently, pound for pound, fish have not quite as much nutritive material as meat, but they are more quickly digested, with the exception of a few kinds in which the per cent. of fat is above six,—*e. g.*, eels, salmon, turbot, herring, and codfish, the fibers of which are coarser than those of the other white fish. Salting and drying fish hardens the fibers and thus preserved fish are not as quickly or as perfectly digested as fresh fish.

Mollusks.—The mollusks most frequently used for food are oysters, clams, mussels, and scallops. They have not a very high nutritive content; oysters, for example, contain about 88.3 per cent. water, but they are valuable for their flavor, and oysters, raw or properly cooked, are quickly and fully digested. The flesh of the other mollusks, especially scallops and mussels, is tougher and therefore not as digestible as oysters. As can be seen in the table, page 340, the mollusks contain glycogen.

Crustaceans.—The crustaceans usually used for food are lobsters, crabs, and shrimps. The fibers of the flesh of the crustaceans, especially lobsters, are coarse and consequently these foods are neither easily nor perfectly digested.

Milk

The milk secreted by any of the mammalian class of animals is suited to the requirements of their young and, as these differ, the quantitative composition of the milk of different animals differs also.

The two milks most frequently compared are human milk and cow's milk, the latter being often substituted for the former in infant feeding.

The percentage composition of average samples of each milk is as follows:

	Water	*Casein-ogen*	*Lact albumin*	*Fat*	*Lactose*	*Ash.*
Human milk	87.41	1.03	1.26	3.78	6.21	0.31
Cow's milk	87.17	3.93	0.53	3.64	4.88	0.71

Digestibility and nutritive value of milk.—As can be seen by comparing the composition of milk with that of other foods (page 340), milk, though a liquid, contains less water and more solid matter than some vegetables, and, usually, it is one of the most perfectly digested foods so that there is little or no waste. The high calcium content of milk makes it a particularly valuable food for children who need calcium for the hardening of their bones.

As shown in the accompanying table, there is more caseinogen and less albumin in cow's milk than in human milk, and this is an important difference when cow's milk is used for infant feeding, since the soluble caseinogen is changed by acid and by the gastric rennin into an insoluble substance, casein, which forms a solid curd, and lact albumin undergoes no such change. When milk is heated above 150° F. the caseinogen will not be clotted by rennin. The reason for this is not known, but it is thought that it is due to change in the nature of the calcium salts by the heat, the calcium salts being, it is supposed, necessary for the change. The reason for this supposition is that caseinogen will not change to casein upon the addition of rennin to milk if the calcium salts have been changed by combining with oxalates. This can be seen by the following experiment.

Experiment 40. Procedure: Into each of three test tubes pour about 5 c. c. of milk; heat that in tube 2 to a temperature of 150° F., and maintain this temperature for a few minutes; add 10 drops of ammonium oxalate solution to the milk in tube 3 and then, to the milk in all three tubes, add about 10 drops of rennin solution. Stand the tubes in a water bath and keep the temperature about 40° C. for 30 minutes.

Only the milk in tube 1 should be affected by the rennin.

Certified milk.—By certified milk is meant milk: (1) that is drawn from cows that have been declared disease-free by an authorized inspector; (2) that is drawn and kept under conditions laid down by the Board of Health; (3) that does not contain a larger number of bacteria than the law of the State permits; and (4) that contains the percentage of fat required by the State law. The details of the law requirements vary in different States. New York State requires that there be less than 30,000 germs per cubic centimeter of milk and that these be non-pathogenic, that the milk contain an average of 4 per cent. fat, that it be kept and sold in sealed, sterilized bottles and sold on the same day that it reaches the city.

To prevent the multiplication of germs in milk it is most important that everything coming in contact with it be absolutely clean and that it be kept at a temperature of or below 40° F.

Pasteurized milk.—By pasteurized milk is meant, milk that is heated sufficiently to destroy non-spore-bearing bacteria. Milk pasteurized by what is known as the *holder* method is brought to a temperature of 145° F., and held there for thirty minutes; that pasteurized by what is known as the *flash* method is gradually brought to a temperature of 160° F. and the temperature maintained for one minute. In both methods the milk must be cooled as quickly as possible, and it is just as important to keep pasteurized milk cold as fresh milk, since spores are not killed at a temperature of 145° F. or even 160° F. and they will develop into bacteria unless the milk is kept at a low temperature.

Experiment 41. Object: To test the comparative effectiveness of the holder and the flash methods of pasteurization.

Procedure: Into each of twelve sterile test tubes put about 5 c. c. of fresh milk, plug the tubes with sterile absorbent cotton, and put six of the tubes into one water bath and six into another. Have the water in both baths cold. Raise that in one to 145° F., and maintain it at this temperature for thirty minutes. Raise the temperature of the water in the other bath to 160° F. And let it remain thus for one minute.

Chill the milk in both tubes as quickly as possible by standing them in cold water and, as soon as they are cool, put them in the ice-box. After twenty-four hours, and every succeeding day, taste the milk in a tube from each test. In which set of tubes did the milk remain sweet longest?

There is a considerable difference of opinion as to whether the digestibility and wholesomeness of milk is increased or decreased by pasteurization. Some authorities contend that pasteurization is beneficial because it kills disease-producing germs and prevents the caseinogen forming a clot in the stomach. Others consider that it is better to use certified milk than pasteurized milk because: (1) they think the change in the salts which prevents the forming of casein is not beneficial; and (2) the bacteria which cause the fermentation of lactose and consequent production of lactic acid are more easily killed than other varieties which they, when present, prevent multiplying. Some of these varieties, often present, cause changes in the proteins of milk that give rise to substances that are injurious, especially to children, but which, unless present in excess, do not advertise their presence, as does lactic acid, by souring or otherwise flavoring the milk.

Sterilized milk.—By this is meant boiled milk. This is not suitable for infant feeding because of the changes caused in the salts by the high temperature and the loss of nutriment in the scum that forms on the top of the milk when it reaches a temperature of about 170° F. The chief constituent of this scum is lact albumin, but it also contains some salts, caseinogen, and fat.

Evaporated and condensed milk and cream.—These are prepared by heating the milk or cream in a vacuum pan until a large portion of the water is evaporated and germs are killed, and then putting the liquid into sterile cans and sealing these at once.

Desiccated milk.—This is milk that has been evaporated to a dry powder. It is prepared for use by adding water to it.

Koumiss, matzoon, artificial buttermilk.—These and similar preparations are fermented milks. In some preparations the fermentation is produced by the use of yeast, in others by that of pure cultures of lactic acid, producing bacteria. In any case, the substance fermented is the glucose, and the curding of the protein is caused by the lactic acid resulting from this fermentation. These milks are quickly digested and are often used by those suffering from impaired digestive power. They are used also in the treatment of intestinal putrefaction, for it is believed that the lactic acid inhibits the action of the putrefactive bacteria always present in the intestine.

Whey.—This is the watery, straw-colored liquid that separates from the curd formed in milk by rennin or acid. Whey consists of water, lact albumin, lactose, and mineral matter.

Cream.—Cream is the name given to the fat which, in the form of an emulsion, rises to the top of milk when it is allowed to stand. The top layers of cream will contain more fat and less protein than those nearer the surface of the milk. Cream is also separated from milk by the use of the centrifugal separator, and that obtained in this way is much richer than that obtained by gravity and it can be better freed from entangled caseinogen. The fat of cream being in the form of an emulsion, is digested more easily than other fats.

Butter and its substitutes.—Butter is made by agitating cream, usually in utensils known as churns, and thereby causing the particles of fat to cling together in masses. The melting point of butter is low and it is thus one of the most easily digested of animal fats. When foods are fried in butter, however, their digestion is interfered with, because fat is not digested until it reaches the intestine, and if carbohydrates and proteins are encased in a hardened fat envelope, the digestive juices of the mouth and stomach cannot attack them properly.

Renovated or process butter is made by melting rancid butter, removing any disagreeable odor and taste by various means, and churning the butter with some new milk to which, usually, cultures of bacteria that will develop an agreeable flavor have been added. The great objection to the use of renovated butter is that the germs which produce rancidity sometimes develop injurious substances in the butter.

Margarin, oleomargarin, and the majority of other artificial butters are made from meat fats with, usually, a little butter or milk to improve the flavor. These substances can be sold at a lower price than good, fresh butter, and though not as good from any dietetic standpoint as fresh butter, they will, when properly prepared, keep better than butter and they are more wholesome than renovated or old butter.

Cheese.—Cheese is made by coagulating the caseinogen of milk. Usually this is done with rennin. As a rule, there is a considerable, though very varying, amount of fat entangled in the casein. As can be seen by studying the table on pages 340–344, cheese contains less water and more protein than the majority of foods. As it is such a concentrated food, it cannot be eaten in large amounts, especially by persons of sedentary habits, and it is very likely to cause gastric disturbances in those whose digestion is impaired, but it has been found that 95 per cent. of the fat and 92 per cent. of the protein are ultimately absorbed by the average healthy adult and that cheese stimulates the secretion of the digestive juices. As many cheeses are cheaper than eggs and meat, cheese is a valuable source of protein food for healthy adults of restricted means. Cooking cheese renders it less digestible since it hardens the casein.

Eggs

Eggs, though lacking carbohydrates, are a particularly valuable food, for they contain easily digested proteins, more lecithin than other foods, and an abundance of mineral matter, especially phosphates and iron, and the iron is in a particularly assimilable form.

The protein of the white is chiefly albumin, that of the yolk is partly albumin, but largely vitellin—a phosphoprotein; part of the fat is in the form of lecithin. The sulphur present in egg, both in the white and yolk, will sometimes give rise to flatulence, because, especially when a person is constipated, the sulphur is likely to unite with hydrogen and form hydrogen sulphid, a light ill-smelling gas. When silver comes in contact with eggs, its surface blackens, due to the union of the sulphur with it and the consequent formation of silver sulphid which remains as a covering on the metal.

Honey

This is a mixture of sugars prepared by bees from substances they extract from the exudations of various plants. Genuine honey consists chiefly of levulose and glucose, with a small amount of sucrose and about 17 per cent. water. The quality and flavor of honey depend largely upon the nature of the flowers from which the bees secured their food material. As honey consists so largely of invert sugars, the greater portion of its substance is ready for absorption and does not undergo digestion.

Plant Foods

Classification.—The plant foods in common use are classified as cereals, vegetables, fruit, nuts, fungi, and sugars.

Cereals

This class of foods includes plants of the grass family such as wheat, rye, oats, barley, rice, and corn. The composition of the cereals varies, but they are all rich in carbohydrates and mineral matter, and cooking does not entail any loss of nutriment.

Wheat.—Wheat, especially in the form of flour, is one of the most universally used of the cereals. There are several varieties of wheat, and the flours and breakfast foods made from them differ because of this and, more especially, as the result of different methods of preparation. The protein of wheat is chiefly gluten, a sticky, tenacious substance which is a compound of gliadin and glutenin. The elasticity of gluten and its power to hold the gas produced in dough by yeast, baking powders, etc., is influenced by the proportions of gliadin and glutenin that it contains—about 65 per cent. of gliadin to 35 per cent. of glutenin is usually preferred for bread flour. Spring wheat (*i. e.*, that sown in spring and maturing in the late summer) contains more gluten and less starch than winter wheat (that sown in the fall and maturing in the early summer) and is preferred for making bread, but winter wheat is the best for pastry.

Oatmeal.—The meal of oats is also much used for food. Oatmeal contains more protein, fat, and mineral matter than other cereals. The protein consists of a substance called *avenin*, which is somewhat similar to the legumin of peas, and a small amount of gliadin. As oatmeal does not contain gluten, it cannot be used for bread-making.

Barley.—Barley contains a relatively large amount of salts, fat, and cellulose, but less protein and digestible carbohydrates than the other cereals. Its salts are soluble in water and hence are easily extracted. This is one reason why barley water is so much used for diluting top milk for infant feeding.

Sprouted and dried barley, generally known as *malt*, contains a ferment called *diastase* which converts starch to sugar.

Corn.—Corn is a widely used cereal both in its natural state and ground into meal and starch. Its protein has not all the characteristics of a true gluten and thus its meal has to be mixed with rye or wheat flour for bread-making. Starch is one of the most valuable products of corn. The nutritive value of corn compares favorably with that of wheat. When eaten in its natural state, corn occasionally causes irritation in the intestines on account of its cellulose exterior, which may resist digestion, but it is stated that analysis of feces has shown that young corn, when properly cooked, is more thoroughly digested than is generally supposed.

Rice.—Rice is very extensively used for food. It is poor in proteins and fat, but rich in starch; it contains a fair amount of mineral matter.

Rye.—The composition of rye is very similar to that of wheat. Its gluten, however, contains more gliadin than does that of wheat and it is therefore more sticky and tenacious and, consequently, bread made with rye flour is darker and less porous than wheat bread.

Substances similar in composition to the cereals, though not belonging to the same botanical classification are: The grain buckwheat; tapioca, which is prepared from the root of the cassava plant; sago, which is prepared from the pith of the sago palm; arrowroot, a starch obtained from the roots of several plants grown in different tropical countries.

Vegetables

Vegetables are often classified as the roots and tubers, green vegetables, and legumes.

The roots and tubers.—The vegetables thus classified are those of which the root or bulb is used for food, *e. g.*, potatoes, parsnips, carrots, beets, turnips, salsify, onions.

These vegetables contain from 75 to 95 per cent. water and they are all relatively poor in protein and in fat, nevertheless they are of value in the diet on account of their flavor and their mineral salts.

Potatoes have a higher nutritive value than the other roots and tubers and are easily digested; they are particularly rich in alkaline potash salts and are thus a valuable adjunct to meat in the diet, the latter being a source of acids in the system. Potatoes, however, contain little iron or calcium salts and therefore, unless mixed with milk or egg, must not form too large a portion of children's vegetable supply.

The carbohydrates of beets, carrots, parsnips, salsify, and turnips are largely cellulose and sugar, and as the latter is soluble in water there is a great loss of nutrient in the cooking of these vegetables.

Green vegetables.—The vegetables often classed as *greens* are those of which the stalks and leaves are used; *e. g.*, cabbage, cauliflower, asparagus, French artichokes, spinach, lettuce. The majority of these vegetables have an even lower nutritive value than the roots and tubers, but they are valuable for their flavor and salts; also they give bulk to the diet, and their cellulose causes a mild irritation of the intestinal tract which stimulates peristalsis and thus helps to prevent constipation.

Legumes.—The principal legumes used as food are peas and beans—of which there are many varieties—and peanuts.

The legumes contain more protein than other vegetable foods, but the protein is in the form of legumin, and this is acted upon mainly by the ferments in the intestine; thus, legumes are not readily digested in the stomach and sometimes, especially when the vegetables are old and the cellulose thus hardened, a considerable portion may escape digestion. Passing peas and dried or Lima beans through a colander after they are cooked, and thus removing the harder cellulose, greatly increases their digestibility. Flatulence is sometimes caused by eating a large amount of beans, because (1) the gas hydrogen sulphid may be formed from the sulphur they contain, and (2) methan gas is sometimes produced, probably by the decomposition of the bean germ by the bacteria that are always present in the intestine.

Eggplant, cucumber, vegetable marrow, pumpkins, tomatoes, from the manner of their growth, are sometimes classed as fruit, though they are used as vegetables. Their nutritive value is not high, but they give bulk and variety to the diet and furnish mineral matter. The cellulose of cucumbers is too difficult of digestion to allow of the use of this plant by invalids and small children.

Fruit

Fruits consist principally of water holding in solution sugars, small amounts of proteins, pectin, and, in some fruits, starch; also they contain acids, essential oils and aromatic oils to which their flavor is due, and cellulose. The cellulose in the immature fruit is insoluble and very indigestible, but it becomes softer and more soluble as the fruit matures. Though the majority of fruits have not a very high food value they are of use for their flavor, for the bulk which they add to the diet, for their mineral matter and their acids. The principal acids of fruit are: malic acid, found more especially in apples, pears, currants, berries, grapes, cherries, pineapples; citric acid, which occurs more especially in lemons, limes, grape fruit, oranges, currants, and gooseberries; tartaric acid, found chiefly in grapes; racemic acid, an acid that resembles tartaric acid, and is found chiefly in grapes. Acetic acid is found in fruits undergoing decay, as it arises from the fermentation of grape sugar, but it is not a constituent of good fruit.

Nuts

By looking at the table on page 344, it will be seen that nuts contain comparatively little water and a high content of protein and fat, and experiments have shown that when nuts are not eaten in large quantities at a time or at the end of a heavy meal, they are much more quickly and thoroughly digested than was previously supposed. Nuts and fresh fruit form a good combination, nuts being deficient in carbohydrate and a very concentrated food, and ripe fruit containing easily digested carbohydrate and being bulky in proportion to the food value of its contents.

Fungi

The fungus in most common use is the mushroom. Judging from the chemical analysis of the mushroom, as shown on page 343, this fungus would seem to have the same food value as the green vegetables, but it has been found that mushrooms are neither thoroughly digested nor absorbed, and this, of course, lessens their nutritive value.

Condiments, Spices, and Flavoring Extracts

By condiments are meant substances, such as the various peppers and salts, used to flavor food.

Condiments, with the exception of sodium chlorid, spices, and other flavoring matter, are obtained from various parts of plants, some being prepared from the stems and leaves (*e. g.*, bay leaf, sage, sweet marjorum), others from the bark (*e. g.*, cinnamon), others from the buds and flowers (*e. g.*, cloves and saffron), some from fruit (*e. g.*, allspice, capsicum, juniper, pepper, vanilla), and still others from the seeds (*e. g.*, anise, caraway, celery seed and salt, coriander, mustard, nutmeg).

Many of these substances owe their properties to the presence of a volatile oil which can be extracted and used separately. One method of extraction is by distilling the oil with steam; oil of cloves, oil of cinnamon, and several others are prepared in this way. Pressure is another method of obtaining oil from plants. The oil of lemon, from which lemon extracts are prepared by dissolving the oil in alcohol, is extracted from lemon rinds by pressure. In some instances the active principle or flavoring matter is extracted from a substance by the use of a solvent, *e. g.*, vanilla extract is made by extracting the active principle of the vanilla bean—*vanillin*—by soaking the bean in dilute alcohol. Artificial, so-called, *vanilla extracts* are now made, which are cheaper than, but not as good as, the true extracts. One variety is made by oxidizing eugenol, a substance in the oil of cloves, with potassium permanganate.

Food value.—Condiments have no food value in themselves, but if not used in excess they help in the digestion of food by stimulating the flow of digestive juices. They do this in two ways: (1) by improving the aroma and flavor of the food, and (2) by a slight irritation of nerve endings in the gastric mucous membrane. If the irritating condiments are used in excess, the irritation they produce may cause congestion of the gastric mucous membrane and consequent digestive disturbances.

Beverages

Caffein and theobromin beverages.—The caffein and theobromin beverages in most common use are coffee, tea, chocolate, and cocoa.

Coffee

Coffee is prepared from the berry of the coffee tree by hulling and drying the berry and, later, roasting and drying the seeds. Sometimes the seeds are glazed in order to prevent a loss of the volatile substance to which the flavor and aroma of coffee are due. The glaze ordinarily used is prepared from sugar, and the pure food laws of the United States prohibit the use of other glazes unless their presence and nature are stated on the label, but some countries permit the use of resin, shellac, and similar substances.

Sources of coffee supply.—True Mocha coffee, which has been always very generally considered one of the best coffees, comes from Arabia, but much of the so-called Mocha now sold comes from the West Indies, Brazil, and other South American countries. Java coffee, from the island of Java, is another favorite coffee, but much of the so-called Java comes from various places and may or may not be good. That coming from Bogota, S. A., is said to be most like the true Java. More than half the coffee supply for the world now comes from Brazil and various localities in the northern part of South America. Other important sources of supply are Central America and the West and East Indies.

Composition of coffee.—Roasted coffee contains 0.6 to 2 per cent. of caffein, a varying amount of caffeol—the volatile oil to which the aroma and flavor are due—and a large amount of caffeo-tannic acid. A cup of coffee solution, as usually prepared, may contain from 1½ to 3 grains of caffein; the amount of caffeo-tannic acid will vary according to the method of preparing the beverage. If the coffee is prepared without boiling, or if it is boiled but a short time and used, or removed from the grounds, at once, it will contain but a small amount of the acid. The caffeol is developed during the roasting of the seeds and, being volatile, it gradually decreases in amount after the seeds are ground. Therefore the flavor of coffee will be stronger if it is not ground until shortly before use, and there will be a great loss of flavor unless the ground coffee is kept in air-tight cans and the coffee-pot kept covered while the beverage is being prepared.

Action of coffee constituents on the system.—The caffein of the coffee acts as a cerebral stimulant. If coffee is not taken in excess, it renders the mind clearer and more alert and it increases the power of concentration; in fact, all the higher intellectual faculties, as reasoning, self-control, judgment, may be stimulated. Through the influence of caffein on motor cells of the nervous system, coffee improves muscle tone, and stimulates the circulation and the respiration; also, it lessens fatigue and may promote a sense of well-being. If coffee is taken in excess, the caffein over-stimulates the nervous system and will lessen the powers of concentration, clear judgment, etc. Caffeol acts as a local irritant in the intestine, thus promoting peristalsis and having a laxative effect. It will, if coffee is taken in excess, cause digestive disturbances, as will also the caffeo-tannic acid. The latter, if much is present in the coffee, retards digestion and absorption. It does not, as does the tannic acid of tea, precipitate alkaloids, albumin, and gelatin; therefore, coffee is useless as an antidote for poisoning by drugs the active principles of which are alkaloids. As caffein stimulates the respiratory center, coffee is often used as an antidote for poisoning by drugs, such as opium, which depress it. Also, as caffein stimulates the vaso-constrictor and motor centers, coffee is sometimes used to counteract the conditions present in collapse, especially that produced by overdoses of drugs which depress the nervous system, as the narcotics, anesthetics, and alcohol.

Caffein—Free Coffees

There are some coffees from which, by heating the berries under special conditions, much of the caffein has been removed. Some of these contain as little as 0.3 per cent. caffein, but it is doubtful if any of them are actually free from it, as the name implies. It is claimed that the other ingredients of these coffees are not altered by the treatment.

Coffee Substitutes

There are various preparations on the market that consist principally of such substances as parched corn, baked wheat, dried peas, bread crust, and the like, which are called *coffee substitutes*. Many of them, however, contain a considerable amount of coffee.

Tea

Tea is made from thea, an evergreen shrub grown chiefly in India, Ceylon, China, and Japan. The finest qualities are made from the young, tender leaves. The chief difference between black and green teas is the smaller amount of tannin in the former. This difference is due to the manner in which the leaves are cured, those for black tea being allowed to undergo a certain amount of fermentation before they are dried.

The chief teas from India and Ceylon are Flowery Pekoe, Orange Pekoe, Pekoe—these are made from the smaller leaves; Souchong, Congou, and Bohea—made from the larger leaves. The principal China teas are Young Hyson and Gunpowder, which are made from the smaller leaves, and Imperial and Twankay, made from the larger leaves. These are green teas. Various teas are often blended and different blends are known by different names. *English breakfast tea*, for example, consists of a mixture of teas.

The principal constituents of tea.—These are: 1 to 4 per cent. caffein, 0.6 per cent. volatile oil—to which the aroma and flavor are due,—and a large amount of tannic acid. Tannic acid precipitates gelatin, albumins, and alkaloids and is strongly astringent. Due to the action of tannic acid upon alkaloids, a strong infusion of tea is useful as an antidote in poisoning by alkaloids; due to the astringency of tannic acid, strong infusions of tea will retard digestion by inhibiting the flow of digestive juices. If tea is not boiled nor the water allowed to remain more than two or three minutes on the leaves, there will not be much tannic acid in the infusion, but, infusions strong enough to tan hide into leather can be made from tea.

Tea contains more caffein than coffee, but, as less of the tea is used per cup, a cup of tea will not contain as much caffein as a cup of coffee. Tea, however, has a more immediate stimulating effect, either because of the nature of its volatile oil or because its absorption is more rapid. If tea is taken in excess or if it is made in a manner to extract a large amount of tannic acid, it will have more injurious effects upon digestion than coffee, owing to the astringent nature of the tannic acid, but, on the other hand, if it is properly made, tea, probably because it contains less extractive matter, is less disturbing to the stomach.

Neither tea nor coffee is nutritive in itself, but the addition of cream and sugar gives them a slight food value. The tannic acid of the tea precipitates the coagulable protein of the milk, but this does not interfere with its digestion by the gastric juice; in fact, it is thought that milk or cream may be desirable additions in that they lessen the irritating action of the acid in the stomach and retard the absorption of caffein.

Chocolate

Chocolate is made from the ripe seeds of the bean of the *theobroma cacao*, a plant which is now grown in many tropical countries. The seeds are fermented, dried, roasted, and deprived of their shells, which are known as *cocoa nibs*.

Chocolate contains from 0.3 to 2 per cent. of theobromin, 10 per cent. of starch, 15 per cent. of vegetable protein, and 30 to 50 per cent. of a fat which is known as *cacao* or *cocoa butter*. There is some tannic acid in the seeds, but this is removed in the fermentation process. The chocolate flavor is developed in the roasting. This flavor is not pleasant, but is made so by the addition of sugar to the chocolate and, as a rule, vanilla or some other flavoring extract.

As shown by the list of its constituents, chocolate, unlike tea and coffee, is nutritious in itself, but, in the stomach, the fat retards both the secretion of the gastric juice and the motor action of the stomach, so that chocolate should not be eaten or drunk in large quantities and it does not make a suitable beverage for invalids.

Theobromin stimulates both the cardiac and voluntary muscles, but has no action, like caffein, on the vaso-constrictor centers and but little stimulating effect upon the brain.

Cocoa

Cocoa is a powder prepared from chocolate by the removal of a portion of its fat by hydraulic pressure. Thus, cocoa contains a smaller per cent. of fat than chocolate (it has 15 to 30 per cent.) and a larger per cent. of theobromin. Considerable extra starch is often mixed with the cocoas sold cheaply and in such cases the percentages of both fat and theobromin are reduced.

When preparing cocoa to drink, it should be cooked for at least five minutes to hydrolize the starch. When made with milk and sweetened with sugar, a cup of cocoa will yield about 250 calories. Thus it has a comparatively high food value for a beverage and it is slightly stimulating, but, even though it has less fat than chocolate, it sometimes interferes with digestion; especially in the case of invalids and of children.

Coco-Cola

Coco-cola is a proprietary preparation that has recently come into use. When it was first put upon the market, it was reported to contain cocaine, but investigation proved the report untrue. The chief constituents are flavoring compounds and caffein. Its action is therefore similar to that of tea and coffee.

Mineral and Carbonated Waters

As stated on page 209, in certain localities the soil and rocks contain so much soluble mineral matter that water flowing through the district absorbs enough of the mineral to give it a decided flavor.

There is so much demand for such waters that they are now bottled at the places where the springs occur and put upon the market.

Both mineral and ordinary drinking waters are often charged with carbon dioxid gas. The carbon dioxid used for the purpose is obtained from various sources. Some of these are: (1) Many of the mineral waters in their natural state contain carbon dioxid, and this is taken from the water at the springs where it arises from the ground, compressed in cylinders, and then used to charge the water when the latter is bottled. (2) Burning some form of carbon, as coke or charcoal. (3) Collecting the by-product (CO_2)

that occurs in the manufacture of lime from calcium carbonate and that arising during the brewing of beer and other fermentative processes.

Unfermented Beverages Made from Fruit Juices

Agreeable and wholesome beverages are prepared by extracting the juice from fruit, sterilizing it, and bottling it in sterile, air-tight bottles.

Fruit syrups are made in the same way plus the addition of sugar before the sterilization.

Imitation fruit syrups are made by combining sugars or saccharin, artificial coloring matter, fruit acids (*e. g.*, tartaric, citric, phosphoric), and flavoring substances prepared from such compounds as chloroform, nitrous ether, benzoic ether, butyric ether, etc., these, when in special combinations, having somewhat the same flavor as the delicate ethers and aldehydes to which the flavor of fruit is due. The syrups used at soda fountains are often the imitation products.

Fruit vinegars are other common flavoring substances used for beverages. They are made in about the same way as the fruit juices, except that the fruit is soaked in vinegar before the juice is extracted and this vinegar is used as well as the juice.

Alcoholic Beverages

The alcoholic beverages in common use are made by fermenting sugar solutions with yeast in the presence of nitrogenous substances. The sugar may be that of a fruit juice, or that produced by the action of the yeast, or ferments secured from other sources, on starch or cellulose.

Alcoholic beverages are usually classed as: wines, distilled liquors or spirits, elixirs, malt liquors.

Wines

Wines are spoken of as red, white, dry, sweet, strong, light, sparkling, fortified.

Red wines are prepared by fermenting the juice of grapes in the presence of their skins. They contain tannic acid and are therefore more astringent than white wines.

White wines are made from the juice of grapes the skins and seeds of which have been removed.

A sweet wine is one that contains free sugar.

A dry wine, one that contains no free sugar, *e. g.*, sherry.

A sparkling wine is one that contains CO_2, *e. g.*, champagne.

A light wine is one that has a low per cent. of alcohol, and a **strong or heavy wine** is one that contains a comparatively large amount of alcohol, but not above 17 per cent., while a **fortified wine** is one that contains a higher per cent. of alcohol than 17 per cent. Since the action of yeast is inhibited when 15 to 17 per cent of alcohol is produced in the fermenting material, to obtain a beverage containing more than 17 per cent. alcohol, the alcohol must be distilled from the fermenting material by fractional distillation—as described on page 63—or else, as is done in the making of fortified wines, alcohol obtained by distillation must be added. Light wines, such as claret and sauterne, contain about 7 to 12 per cent. alcohol. Fortified wines, such as madeira, sherry, and port, contain about 17 to 25 per cent. alcohol.

Apple and pear cider and similar beverages are often classed with wines since they are prepared by the fermentation of fruit sugar. They usually contain from 0.50 to 1 per cent. alcohol.

Distilled Liquors or Spirits

These are made by distilling different kinds of fermented liquors. Examples are: whiskey, gin, brandy, and rum.

Whiskey (*spiritus frumenti*) is prepared by distilling the mash of fermented grain, such as corn, rye, wheat, barley. To be good, whiskey must be not less than four years old. It contains 44 to 55 per cent. by volume of ethyl alcohol.

Gin, sometimes called *compound spirits of juniper*, is made by redistilling the distillate obtained in making whiskey with juniper berries. It contains from 60 to 70 per cent. alcohol and some volatile oil of juniper berries. The latter gives it diuretic properties.

Brandy (*spiritus vini gallici*) is obtained by the distillation of the fermented juice of grapes. It contains 46 to 55 per cent. by volume of ethyl alcohol.

Rum is the distillate from fermented molasses. It varies in strength, but usually contains from 45 to 55 per cent. alcohol.

The Elixirs

These are aromatic alcoholic mixtures containing varying, but large, percentages of alcohol, flavoring substances, and sugar. Elixirs are often colored red, blue, yellow, or green with vegetable colors.

The elixirs include the various liqueurs, cordials.

crêmes, etc. Some of the best known beverages of this class are: **Absinth**, which is prepared from spirit flavored with wormwood and various aromatic substances and colored green with extracts from the green leaves of wormwood, hyssop, and mint. **Chartreuse**, made from alcohol distilled with angelica, hyssop, nutmeg, and similar substances. **Crême de menthe**, made by distilling a fermented decoction of mint, cinnamon, sage, and other spices; it is sweetened with sugar and colored with chlorophyl. **Curaçao**, made by macerating and distilling the dried peel of a variety of orange that grows in Curaçao (Dutch West India Islands) with alcohol. **Maraschino**—genuine maraschino is made from Marasca cherries which are crushed with their pits, mixed with honey, fermented and distilled, but much of that sold is made by distilling the products obtained by crushing and macerating peach pits, wild cherries, cherry leaves, orris root, and other flavoring substances with alcohol. **Vermouth**, made in France and Italy from white wines and various flavoring extracts.

Malt Liquors

The malt liquors are made from starchy substances, usually grain. The grain is ground and boiled with water; barley malt is added after the mash has been reduced to the desired temperature, and as the barley contains the ferment diastase, it changes the starch to dextrin and some of the dextrin to maltose and glucose. Hops also are added and they yield a bitter principle and hypnotic substance to the liquid. When this part of the process is completed, the liquid is removed by filtration and yeast is added to the filtrate. After the fermentation caused by the yeast has been carried to the desired degree, the yeast is killed by heat. As the fermentation is always stopped before all the sugars are destroyed, the malt liquors contain varying amounts of free sugar. Also, they contain the various products of the fermentation of sugars, viz., alcohol (3 to 7 per cent.), acids and carbon dioxid, and the bitter principle derived from the hops.

The principal malt liquors are ale, beer, porter, and stout. The ordinary varieties of ale and beer are very similar; the slight difference that exists being due chiefly to the different yeasts that are used: beer being fermented by *bottom yeast* (yeast that sinks), and ale by top yeast (yeast which floats). True lager beer—so called from the German *lager = a storehouse*—is beer that is stored in a cool place for several months before use. Porter and stout are ales of which the malt, before it was fermented, was roasted until it changed to caramel. Stout is the richer and stronger of the two.

Beers and ales made from roots, herbs, etc.—Such beverages usually consist of sugar solutions flavored with essences consisting of extracts from special roots, etc. (as ginger, sarsaparilla), and fermented with yeast. They contain about 1 per cent. alcohol.

Action of alcohol on the system.—Alcohol depresses the higher brain centers, *i. e.*, those controlling reasoning, judgment, etc., but, unless the alcohol is taken in large amounts, the centers concerned with imagination, emotion, etc., are not particularly depressed, especially when the drinker is where there are bright lights and congenial company; in fact, the depression of the higher centers often allows the others freer play. This effect is more pronounced when gas-containing beverages, as champagne, are used, on account of the stimulating effect of the CO_2. But if alcoholic beverages are taken in excess, and sometimes even when taken in small amounts in the absence of company or other stimulant, they are narcotic. From the depression of the cerebrum by alcohol, after heavy drinking, sexual desires are less under restraint than usual, hence alcohol has been always an important factor in the *white-slave trade.* Alcohol depresses the vaso-constrictor centers and, consequently, it causes the blood-vessels in the skin to become dilated; thus small doses may temporarily improve the superficial circulation and lessen internal congestion. Large doses of alcohol, however, depress this and centers regulating other vital activities to such an extent that, if the depression is constantly repeated, harmful changes will be caused in the body tissues which will interfere with the proper functioning of the body organs and result in mental deterioration. Alcoholic beverages, by stimulating the taste-buds and olfactory nerves act as appetizers and induce a psychic secretion of digestive juices. Beverages containing less than 20 per cent. alcohol do not seem to affect the digestive ferments, and they hasten the absorption of other substances, such as digestive products, water, and drugs, but strong solutions will precipitate the proteins of food and, to some extent, the pepsin of the gastric juice, and they retard digestion; also, they tend to produce a coat of thick mucus on the mucous membrane of the stomach which interferes with absorption.

As to whether alcohol is to be considered a food or not is still an unsettled question. An accepted definition of a food is *a substance whose dominant property* *in the body is to build up the tissues or to yield energy,* and though every gram of alcohol oxidized in the body will yield 7 calories, this can be hardly called its dominant property. Also, experiments have shown that, though alcohol is easily oxidized in the body and thus spares food products and body tissue from oxidation, it is not used in the body for tissue building. The malt beverages and others containing sugar have, of course, a real food value on account of their sugar. This sugar, however, is particularly likely to undergo fermentation in the stomach and cause flatulence.

AVERAGE COMPOSITION OF COMMON AMERICAN FOOD PRODUCTS[1]

Food materials (as purchased)	Refuse	Water	Protein	Fat	Carbohydrates	Ash	Fuel value per pound
ANIMAL FOOD	*Per cent.*	*Per cent.*	*Per cent.*	*Per cent.*	*Per cent.*	*Per cent.*	*Calories*
Beef, fresh:							
Chuck ribs	16.3	52.6	15.5	15.0	.....	0.8	910
Flank	10.2	54.0	17.0	19.0	.....	.7	1,105
Loin	13.3	52.5	16.1	17.5	.....	.9	1,025
Porterhouse steak	12.7	52.4	19.1	17.9	.....	.8	1,100
Sirloin steak	12.8	54.0	16.5	16.1	.....	.9	975
Neck	27.6	45.9	14.5	11.9	.....	.7	1,165
Ribs	20.8	43.8	13.9	21.2	.....	.7	1,135
Rib rolls	.....	63.9	19.3	16.7	.....	.9	1,055
Round	7.2	60.7	19.0	12.8	.....	1.0	890
Rump	20.7	45.0	13.8	20.2	.....	.7	1,090
Shank, fore	36.9	42.9	12.8	7.3	.....	.6	545
Shoulder and clod	16.4	56.8	16.4	9.8	.....	.9	715
Fore quarter	18.7	49.1	14.5	17.5	.....	.7	995
Hind quarter	15.7	50.4	15.4	18.3	.....	.7	1,045

[1] Reprinted from *Farmers' Bulletin*, No. 142.

[To translate the percentages given in this table into grams, divide the figures representing the percentages by 100 (*i. e.*, move the decimal point two places toward the left), and to find the amount of heat (calories) that each food will yield, multiply the number of grams of protein by 4, of fat by 9, and of carbohydrates by 4. To find the amount of protein, fat, etc., in one ounce, multiply the number of grams of protein, etc., by 28.35 (which is the number of grams in one ounce), and to find the amount in one pound, multiply by 453.60.]

AVERAGE COMPOSITION OF COMMON AMERICAN FOOD PRODUCTS—*Continued*

Food materials (as purchased)	Refuse	Water	Protein	Fat	Carbohydrates	Ash	Fuel value per pound
ANIMAL FOOD—*Continued*	*Per cent.*	*Per cent.*	*Per cent.*	*Per cent.*	*Per cent.*	*Per cent.*	*Calories*
Beef, corned, canned, pickled, and dried:							
Corned beef	8.4	49.2	14.3	23.8	.....	4.6	1,245
Tongue, pickled	6.0	58.9	11.9	19.2	.....	4.3	1,010
Dried, salted, and smoked	4.7	53.7	26.4	6.9	.....	8.9	790
Canned boiled beef	.....	51.8	25.5	22.5	.....	1.3	1,410
Canned corned beef	.....	51.8	26.3	18.7	.....	4.0	1,270
Veal:							
Breast	21.3	52.0	15.4	11.0	.....	.8	745
Leg	14.2	60.1	15.5	7.9	.....	.9	625
Leg cutlets	3.4	68.3	20.1	7.5	.....	1.0	695
Fore quarter	24.5	54.2	15.1	6.0	.....	.7	535
Hind quarter	20.7	56.2	16.2	6.6	.....	.8	580
Mutton:							
Flank	9.9	39.0	13.8	36.9	.....	.6	1,770
Leg, hind	18.4	51.2	15.1	14.7	.....	.8	890
Loin chops	16.0	42.0	13.5	28.3	.....	.7	1,415
Fore quarter	21.2	41.6	12.3	24.5	.....	.7	1,235
Hind quarter, without tallow	17.2	45.4	13.8	23.2	.....	.7	1,210
Lamb:							
Breast	19.1	45.5	15.4	19.1	.....	.8	1,075
Leg, hind	17.4	52.9	15.9	13.6	.....	.9	860
Pork, fresh:							
Ham	10.7	48.0	13.5	25.9	.....	.8	1,320
Loin chops	19.7	41.8	13.4	24.2	.....	.8	1,245
Shoulder	12.4	44.9	12.0	29.8	.....	.7	1,450
Tenderloin	.....	66.5	18.9	13.0	.....	1.0	895
Pork, salted, cured, and pickled:							
Ham, smoked	13.6	34.8	14.2	33.4	.....	4.2	1,635
Shoulder, smoked	18.2	36.8	13.0	26.6	.....	5.5	1,335
Salt pork	.....	7.9	1.9	86.2	.....	3.9	3.555
Bacon, smoked	7.7	17.4	9.1	62.2	.....	4.1	2,715
Sausage:							
Bologna	3.3	55.2	18.2	19.7	.....	3.8	1,155
Pork	.....	39.8	13.0	44.2	1.1	2.2	2.075
Frankfort	.....	57.2	19.6	18.6	1.1	3.4	1,155
Soups:							
Celery, cream of	.....	88.6	2.1	2.8	5.0	1.5	235
Beef	.....	92.9	4.4	.4	1.1	1.2	120
Meat stew	.....	84.5	4.6	4.3	5.5	1.1	365
Tomato	.....	90.0	1.8	1.1	5.6	1.5	185
Poultry:							
Chicken, broilers	41.6	43.7	12.8	1.4	.....	.7	305
Fowls	25.9	47.1	13.7	12.3	.....	.7	765
Goose	17.6	38.5	13.4	29.8	.....	.7	1,475
Turkey	22.7	42.4	16.1	18.4	.....	.8	1,060
Fish:							
Cod, dressed	29.9	58.5	11.1	.2	.....	.8	220
Halibut, steaks or sections	17.7	61.9	15.3	4.4	.....	.9	475
Mackerel, whole	44.7	40.4	10.2	4.2	.....	.7	370
Perch, yellow, dressed	35.1	50.7	12.8	.7	.....	.9	275
Shad, whole	50.1	35.2	9.4	4.8	.....	.7	380
Shad, roe	.....	71.2	20.9	3.8	2.6	1.5	600
Fish, preserved:							
Cod, salt	24.9	40.2	16.0	.4	.....	18.5	325
Herring, smoked	44.4	19.2	20.5	8.8	.....	7.4	755

AVERAGE COMPOSITION OF COMMON AMERICAN FOOD PRODUCTS—*Continued*

Food materials (as purchased)	Refuse	Water	Protein	Fat	Carbohydrates	Ash	Fuel value per pound
ANIMAL FOOD—*Continued*	*Per cent.*	*Per cent.*	*Per cent.*	*Per cent.*	*Per cent.*	*Per cent.*	*Calories*
Fish, canned:							
Salmon	.....	63.5	21.8	12.1	.....	2.6	915
Sardines	[a]5.0	53.6	23.7	12.1	.....	5.3	950
Shellfish:							
Oysters, "solids"	.....	88.3	6.0	1.3	3.3	1.1	225
Clams	.....	80.8	10.6	1.1	5.2	2.3	340
Crabs	52.4	36.7	7.9	.9	.6	1.5	200
Lobsters	61.7	30.7	5.9	.7	.2	.8	145
Eggs: Hens' eggs	[b]11.2	65.5	13.1	9.3	.....	0.9	635
Dairy products, etc:							
Butter	.....	11.0	1.0	85.0	.....	3.0	3,410
Whole milk	.....	87.0	3.3	4.0	5.0	.7	310
Skim milk	.....	90.5	3.4	.3	5.1	.7	165
Buttermilk	.....	91.0	3.0	.5	4.8	.7	160
Condensed milk	.....	26.9	8.8	8.3	54.1	1.9	1,430
Cream	.....	74.0	2.5	18.5	4.5	.5	865
Cheese, Cheddar	.....	27.4	27.7	36.8	4.1	4.0	2,075
Cheese, full cream	.....	34.2	25.9	33.7	2.4	3.8	1,885
VEGETABLE FOOD							
Flour, meal, etc.:							
Entire-wheat flour	.....	11.4	13.8	1.9	71.9	1.0	1,650
Graham flour	.....	11.3	13.3	2.2	71.4	1.8	1,645
Wheat flour, patent roller process—							
High-grade and medium	.....	12.0	11.4	1.0	75.1	.5	1,635
Low-grade	.....	12.0	14.0	1.9	71.2	.9	1,640
Macaroni, vermicelli, etc.	.....	10.3	13.4	.9	74.1	1.3	1,645
Wheat breakfast food	.....	9.6	12.1	1.8	75.2	1.3	1,680
Buckwheat flour	.....	13.6	6.4	1.2	77.9	.9	1,605
Rye flour	.....	12.9	6.8	.9	78.7	.7	1,620
Corn meal	.....	12.5	9.2	1.9	75.4	1.0	1,635
Oat breakfast food	.....	7.7	16.7	7.3	66.2	2.1	1,800
Rice	.....	12.3	8.0	.3	79.0	.4	1,620
Tapioca	.....	11.4	.4	.1	88.0	.1	1,650
Starch	.....	.....	.....	.....	90.0	.....	1,675
Bread, pastry, etc.							
White bread	.....	35.3	9.2	1.3	53.1	1.1	1,200
Brown bread	.....	43.6	5.4	1.8	47.1	2.1	1,040
Graham bread	.....	35.7	8.9	1.8	52.1	1.5	1,195
Whole-wheat bread	.....	38.4	9.7	.9	49.7	1.3	1,130
Rye bread	.....	35.7	9.0	.6	53.2	1.5	1,170
Cake	.....	19.9	6.3	9.0	63.3	1.5	1,630
Cream crackers	.....	6.8	9.7	12.1	69.7	1.7	1,925
Oyster crackers	.....	4.8	11.3	10.5	70.5	2.9	1,910
Soda crackers	.....	5.9	9.8	9.1	73.1	2.1	1,875
Sugars, etc.:							
Molasses	.....	.....	.....	.....	70.0	.....	1,225
Candy [c]	.....	.....	.....	.....	96.0	.....	1,680
Honey	.....	.....	.....	.....	81.0	.....	1,420
Sugar granulated	.....	.....	.....	.....	100.0	.....	1,750
Maple sirup	.....	.....	.....	.....	71.4	.....	1,250

[a] Refuse, oil. [b] Refuse, shell. [c] Plain confectionery not containing nuts, fruit, or chocolate.

AVERAGE COMPOSITION OF COMMON AMERICAN FOOD PRODUCTS—*Continued*

Food materials (as purchased)	Refuse	Water	Protein	Fat	Carbohydrates	Ash	Fuel value per pound
VEGETABLE FOOD—*Continued*	*Per cent.*	*Per cent.*	*Per cent.*	*Per cent.*	*Per cent.*	*Per cent.*	*Calories*
Vegetables *a*:							
Beans, dried	......	12.6	22.5	1.8	59.6	3.5	1,530
Beans, Lima, shelled	......	68.5	7.1	.7	22.0	1.7	540
Beans, string	7.0	83.0	2.1	.3	6.9	.7	170
Beets	20.0	70.0	1.1	.1	7.7	.9	160
Cabbage	15.0	77.7	1.4	.2	4.8	.9	115
Celery	20.0	75.6	.9	.1	2.6	.8	65
Corn, green (sweet), edible portion	......	75.4	3.1	1.1	19.7	.7	440
Cucumbers	15.0	81.1	.7	.2	2.6	.4	65
Lettuce	15.0	80.5	1.0	.2	2.5	.8	65
Mushrooms	......	88.1	3.5	.4	6.8	1.2	185
Onions	10.0	78.9	1.4	.3	8.9	.5	190
Parsnips	20.0	66.4	1.3	.4	10.8	1.1	230
Peas (*Pisum sativum*), dried	......	9.5	24.6	1.0	62.0	2.9	1,565
Peas (*Pisum sativum*), shelled	......	74.6	7.0	.5	16.9	1.0	440
Cowpeas, dried	......	13.0	21.4	1.4	60.8	3.4	1,505
Potatoes	20.0	62.6	1.8	.1	14.7	.8	295
Rhubarb	40.0	56.6	.4	.4	2.2	.4	60
Sweet potatoes	20.0	55.2	1.4	.6	21.9	.9	440
Spinach	......	92.3	2.1	.3	3.2	2.1	95
Squash	50.0	44.2	.7	.2	4.5	.4	100
Tomatoes	......	94.3	.9	.4	3.9	.5	100
Turnips	30.0	62.7	.9	.1	5.7	.6	120
Vegetables, canned:							
Baked beans	......	68.9	6.9	2.5	19.6	2.1	555
Peas (*Pisum sativum*), green	......	85.3	3.6	.2	9.8	1.1	235
Corn, green	......	76.1	3.8	1.2	19.0	.9	430
Succotash	......	75.9	3.6	1.0	18.6	.9	425
Tomatoes	......	94.0	1.2	.2	4.0	.6	95
Fruits, berries, etc., fresh *b*:							
Apples	25.0	63.3	0.3	0.3	10.8	0.3	190
Bananas	35.0	48.9	.8	.4	14.3	.6	260
Grapes	25.0	58.0	1.0	1.2	14.4	.4	295
Lemons	30.0	62.5	.7	.5	5.9	.4	125
Muskmelons	50.0	44.8	.3	.....	4.6	.3	80
Oranges	27.0	63.4	.6	.1	8.5	.4	150
Pears	10.0	76.0	.5	.4	12.7	.4	230
Persimmons, edible portion	.....	66.1	.8	.7	31.5	.9	550
Raspberries	.....	85.8	1.0	.....	12.6	.6	220

[a] Such vegetables as potatoes, squash, beets, etc., have a certain amount of inedible material, skin, seeds, etc. The amount varies with the method of preparing the vegetables, and cannot be accurately estimated. The figures given for refuse of vegetables, fruits, etc., are assumed to represent approximately the amount of refuse in these foods as ordinarily prepared.

[b] Fruits contain a certain proportion of inedible materials, as skin, seeds, etc., which are properly classed as refuse. In some fruits, as oranges and prunes, the amount rejected in eating is practically the same as refuse. In others, as apples and pears, more or less of the edible material is ordinarily rejected with the skin and seeds and other inedible portions. The edible material which is thus thrown away, and should properly be classed with the waste, is here classed with the refuse. The figures for refuse here given represent, as nearly as can be ascertained, the quantities ordinarily rejected.

AVERAGE COMPOSITION OF COMMON AMERICAN FOOD PRODUCTS—*Continued*

Food materials as (purchased)	Refuse	Water	Protein	Fat	Carbohydrates	Ash	Fuel value per pound
VEGETABLE FOOD—*Continued*	*Per cent.*	*Per cent.*	*Per cent.*	*Per cent.*	*Per cent.*	*Per cent.*	*Calories*
Fruits, berries, etc., fresh [a]:							
Strawberries	5.0	85.9	.9	.6	7.0	.6	150
Watermelons	59.4	37.5	.2	.1	2.7	.1	50
Fruits, dried:							
Apples	.....	28.1	1.6	2.2	66.1	2.0	1,185
Apricots	.....	29.4	4.7	1.0	62.5	2.4	1,125
Dates	10.0	13.8	1.9	2.5	70.6	1.2	1,275
Figs	.....	18.8	4.3	.3	74.2	2.4	1,280
Raisins	10.0	13.1	2.3	3.0	68.5	3.1	1,265
Nuts:							
Almonds	45.0	2.7	11.5	30.2	9.5	1.1	1,515
Brazil nuts	49.6	2.6	8.6	33.7	3.5	2.0	1,485
Butternuts	86.4	.6	3.8	8.3	.5	.4	385
Chestnuts, fresh	16.0	37.8	5.2	4.5	35.4	1.1	915
Chestnuts, dried	24.0	4.5	8.1	5.3	56.4	1.7	1,385
Cocoanuts	[b]48.8	7.2	2.9	25.9	14.3	.9	1,295
Cocoanut, prepared	.....	3.5	6.3	57.4	31.5	1.3	2,865
Filberts	52.1	1.8	7.5	31.3	6.2	1.1	1,430
Hickory nuts	62.2	1.4	5.8	25.5	4.3	.8	1,145
Pecans, polished	53.2	1.4	5.2	33.3	6.2	.7	1,465
Peanuts	24.5	6.9	19.5	29.1	18.5	1.5	1,775
Piñon (*Pinus edulis*)	40.6	2.0	8.7	36.8	10.2	1.7	1,730
Walnuts, black	74.1	.6	7.2	14.6	3.0	.5	730
Walnuts, English	58.1	1.0	6.9	26.6	6.8	.6	1,250
Miscellaneous:							
Chocolate	.....	5.9	12.9	48.7	30.3	2.2	2,625
Cocoa, powdered	.....	4.6	21.6	28.9	37.7	7.2	2,160
Cereal coffee, infusion (1 part boiled in 20 parts water)[c]	.....	98.2	.2	.....	1.4	.2	30

[a] *Ibid.*

[b] Milk and shell.

[c] The average of five analyses of cereal coffee grain is: Water 6.2, protein 13.3, fat 3.4, carbohydrates 72.6, and ash 4.5 per cent. Only a portion of the nutrients, however, enters into the infusion. The average in the table represents the available nutrients in the beverage. Infusions of genuine coffee and of tea like the above contain practically no nutrients.

CHAPTER XX

THE CHEMISTRY OF COOKING

The Action of Heat, Acids, and Sodium Chlorid on Protein—Solubilities of Proteins—The Changes Caused in Meats by Cooking—The Action of Heat and Acids on Starches and Sugars—Solubilities of the Carbohydrates of Food—Comparative Thickening Power of Different Starches—Methods of Making Doughs and Batters Light—The Nature and Action of Pectin.

A KNOWLEDGE of some of the more important chemical and physical changes that occur in food as the result of cooking is a valuable aid in preventing mistakes which greatly lessen the digestibility of food and, consequently, its nutritive value; for food is of no use to the body until it has been absorbed and it cannot be absorbed until it has been digested. Also, as will be seen in the following pages, mistakes in cooking entail a loss of food in other ways.

Protein

The Effect of Heat, Acids, and Sodium Chlorid on Protein

Experiment 42. Object: (*a*) To discover the temperatures at which the albumin of milk and egg, the vitellin of egg, and the proteins of meat coagulate; (*b*) to see the difference in the nature of egg-white when it is coagulated slowly and rapidly.

Procedure: (*a*) Into separate large test-tubes put (1) milk, (2) egg-white, (3) egg-yolk, (4) a few small pieces of meat, and about 10 c. c. of water to keep the meat from sticking and hardening to the glass. Place the four tubes in a beaker of water and set this over a low flame. Watch carefully and test the temperature of the contents of each tube as soon as it begins to coagulate and every few minutes as coagulation increases. Notice particularly the difference in the temperature and the time required for the coagulation of the white and yolk of the egg.

(*b*) Pour some egg-white into a beaker ¼ full of boiling water and let this boil for three minutes. Put some egg-white into a test tube, stand this in a beaker containing cold water, and bring this slowly, taking ten minutes to do so, to boiling temperature. Compare the nature of the two coagula.

How would you cook a so-called *soft-boiled* egg?

How would you poach an egg?

Experiment 43. Object: To test the effect of salt and organic acids on coagulation.

Procedure: Into each of seven test tubes put about 3 c. c. of egg-white. To tube 4 add ½ c. c. of 10% NaCl solution; to tube 5 add 1. c. c. of 10% NaCl solution; to tube 6 add 1 drop of acetic acid 1%; to tube 7 add 5 drops of acetic acid 1%. Put the seven tubes into a water bath and keep the water just below boiling point for five minutes. (Notice the difference in the toughness of the egg heated alone and that of the egg to which the acid and the salt were added.) At the end of five minutes add: To tube 1, 2 drops of 1% acetic acid; to tube 2, ½ c. c. of 10% NaCl solution; to tube 3, 1 c. c. of 10% NaCl; to tube 4, 2 drops of 1% acetic acid; to tube 5, 2 drops of 1% acetic acid; to tube 6, 1 c. c. of 10% NaCl and the same to tube 7. Replace the tubes in the water bath and leave them for five minutes. Watch their contents carefully to detect the changes that occur.

As the result of these experiments, state your opinion as to the desirability of adding salt to the water in which eggs are cooked.

Would you recommend the addition of vinegar to the water?

Would acid, as vinegar or lemon juice, poured over tough meat have any desirable effect?

Will salt rubbed over the surface of meat before roasting help to retain the extractives in the meat? Why? (If in doubt of the correct answer, put a few small pieces of raw meat into a beaker containing water and let this stand for half an hour or more and, in a second beaker, do likewise with some meat you have heated until the outside has been well browned. Judging from the color of the water in the two beakers, decide whether coagulating the protein on the outside of the meat tends to retain the extractives.)

What will be the result of cooking eggs and meat at a high temperature?

From these experiments, it can be seen that if protein is heated in a slightly acid solution or a neutral salt solution, coagulation is accelerated, but that larger amounts of acids (and the same is true of alkalies) prevent coagulation.

Result of cooking on the protein of meat.—Both dry and moist heat coagulate the proteins of meat and thus harden its tissue, but moist heat disintegrates the fibers and softens the connective tissue and thus renders the meat easier to masticate. Thus, cuts of meat that would be very tough if broiled or roasted can be made tender by cooking them in water at a temperature slightly below boiling point. Searing the outside of meat by exposure to a high temperature will prevent the escape of extractive; but, if the high temperature is maintained, the meat will be toughened, and, if the surface of the meat is hardened, the heat will not penetrate. It has been found that roasts are as quickly—and much more uniformly—cooked at 175° C. as at 200° C.

Solubilities of some of the common food proteins.—These are as follows:

Albumins are soluble in water, dilute acids, dilute alkalies, and sodium chlorid solutions.

Globulins also are soluble in dilute acid, alkali, and neutral salt solutions, but insoluble in concentrated sodium chlorid solutions and in water.

Caseinogen, the important phosphoprotein of milk, is insoluble in water, dilute acids, and salts, but dissolves in alkalies to form a salt-like body. It is precipitated by dilute acids, but dissolves in strong acids. It does not coagulate on boiling. It is clotted by rennin. This is due to the conversion of the caseinogen into an insoluble protein called casein. The calcium salts of the milk are necessary for this change as was seen by Experiment 40.

The curd that forms when the rennin acts upon the caseinogen entangles the greater part of the fat of the milk so that the fluid portion which separates from the curd—the whey—contains only the salts, lactose, and lactalbumin. The presence of fat in the milk rather interferes with the forming of a hard curd and, therefore, when milk is curded in order to obtain whey, it is imperative that skimmed milk be used.

Legumin, the chief protein of the legumes, combines with the lime in hard water and is thus rendered very insoluble. Consequently the legumes must be cooked in soft water or else sodium bicarbonate must be added to the water.

Gluten is insoluble in water, but one of its constituents—glutenin—is soluble in very dilute acid and alkaline solutions and the other constituent—gliadin—is soluble in 70 per cent. alcohol.

Nucleoproteins are soluble in dilute alkaline solutions, but insoluble in water and dilute salt and acid solutions.

Gelatin is soluble in warm water and in warm acid, alkaline, and salt solutions, but insoluble in all of these liquids when cold.

Derived proteins and **extractives** are soluble in water, dilute acid, alkaline, and salt solutions.

Mention all the practical applications you can think of to which knowledge of the coagulation and solubilities of proteins can be put in cooking and in the cleaning of cooking utensils.

Are there more albumins or globulins in the meat? (See page 270)

Why is the salt solution used for *corning* beef concentrated?

How will you roast meat?

How will you make a stew?

How make broth?

Give reasons for all procedures.

Carbohydrates

Experiments 44 to 46. Object: To see which of the common carbohydrates are soluble in water and the effect of letting vegetables such as potatoes stand in water.

Experiment 44. Cut a portion of a firm potato into thin strips and let these stand in a beaker of water for at least half an hour. Do likewise with a wilted potato (use a separate beaker).

Experiment 45. Put a little bread into a beaker of water and in a separate beaker a few strips of dry toast and water.

Experiment 46. Put a few strips of cabbage in (1) a beaker containing water, (2) a beaker containing vinegar.

(Record the result of these experiments before reading further.)

The white substance that collected in the bottom of the beakers in which the pieces of potato stood is starch; therefore Experiment 44 shows that starch is insoluble in water, but that it will be extracted from potatoes if they are allowed to stand for any length of time in water. It will be noticed that the wilted potato did not part with its starch to anything like the extent that the firm one did and also that the condition of the wilted potato was improved by standing in water. Nearly all vegetables, soon after they are gathered, lose water by evaporation; this water must be replaced before the vegetable is cooked or served, or it will be neither tender nor crisp.

Since such vegetables as potatoes lose starch while in water, they must lose it during cooking and, as shown in Experiment 45, the loss is greater as the starch becomes dextrinized[1] by heat and when the carbohydrate present is a sugar, as in parsnips, etc. There will be less loss of nutrient from potatoes if they are

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[1] The brown of the toast is dextrin.

cooked before they are peeled and, when baked, a potato loses practically none of its constituents except water.

In Experiment 45, the water in which the bread was soaked remained unclouded, but that in which the toast stood very soon became colored, showing that dextrin is soluble in water. That sugar is soluble in water is seen too often to need demonstrating.

Experiment 46 shows the difference of the action of water and acid on cellulose, the acid dissolving the cellulose to a slight extent. This is why the addition of vinegar to cabbage is thought to aid in its digestion.

To summarize, celluloses and starches are insoluble in water, dextrin and sugars are soluble.

Experiment 47. Object: To show the effect of heat and water on starch.

Procedure: (*a*) Put 1 tablespoonful of starch or flour into a beaker and pour over it 1 cup of boiling water. Stir slightly. Put the beaker over the flame and boil the contents for two or three minutes. Remove it from the flame. Take out some of the larger lumps and break them open.

Why was the interior of the lump so little affected by the heat?

(*b*) Repeat (*a*), but before adding the water, mix thoroughly two tablespoonfuls of sugar with the starch.

(*c*) Repeat (*a*), but blend the flour to a smooth paste with cold water before adding the boiling water.

(*d*) Repeat (*a*), but mix the flour with melted butter before adding the boiling water.

Why did not the starch lump in (*b*), (*c*), (*d*)?

Starch grains are enclosed in a comparatively soft cellulose sac. This, when heated to boiling temperature, is softened, and the grains swell to many times their original size, and, unless they are well separated from each other in some way as by cold water, sugar, or fat, they will stick together, forming lumps. The grains of cereals that have a comparatively thick cellulose sac can be prevented from lumping, even though they are not mixed with sugar, etc., if the water is boiling violently when the cereal is put into it and kept boiling during the process, the bubbling of the water preventing adhesion.

When a starch paste is made with water below boiling temperature, the covering of the grain is unbroken, though the grain swells considerably, and when such a paste stands, the starch grains sink to the bottom.

The comparative thickening power, transparency, and flavor of starches.—Some starch grains swell much more than others and some will make a more transparent paste and have a more delicate flavor than others. This can be seen by the following experiment.

Experiment 48. Into separate beakers, put 2 grains of each of the following starches: Arrowroot, corn, potato, rice and wheat starches, tapioca, dextrin. To each add slowly, 40 c. c. of cold water, bring to boiling point and boil for five minutes; stir constantly. Compare the different mixtures as to consistency, clearness, and flavor. Boil the pastes for five minutes longer and compare them again.

Which forms the clearest paste?

Of which paste is the flavor improved by the longer cooking?

If, when following the directions of a recipe, you substitute the starch specified for another, will you use the same quantity?

Which of the starches used has the greatest thickening power and which the least?

Experiment 49. Object: To study the effect of heat and heat plus acid on starch.

Procedure: Heat some starch in an evaporating dish over a low flame until it browns, dissolve ¼ of a teaspoonful of this in water. Make a starch paste, using ½ teaspoonful of starch to ½ cup of water. Divide this into two equal portions, (*b*) and (*c*). To (*c*) add 1 c. c. of HCl and boil for ten minutes. Divide each of these three mixtures into two portions and test one with iodin and one with Fehling's solution. It will be noticed that the amount of starch used in each case is such that there will be an equal amount in each of the mixtures tested with Fehling's. The reason why this is necessary can be seen on page 398. For the same reason, it is necessary that exactly the same amount of Fehling's solution be used in each case.

Into what was the starch changed by heat?

Into what was it changed by heat and acid?

If unable to answer these questions, see page 398.

Experiment 50. Object: To show the effect of heat and acid on sugar.

Procedure: Make a sugar solution by dissolving one teaspoonful of sugar in 30 c. c. of water. Divide this into three equal portions. To the portion in tube 1 add 5 c. c. of lemon juice; boil this and the solution in tube 2 for fifteen minutes. Divide the three solutions into two parts and test for cane sugar and for glucose as described on page 294.

It will be seen that boiling reduces cane sugar to glucose, but that the reduction is more quickly and thoroughly accomplished when acid is present. Glucose, it will be remembered, is only about half as sweet as cane sugar.

Why is sugar, when possible to avoid it, not added in cooking until the cooking is nearly completed, especially when there is acid in the food?

Will you need more sugar for a baked or for a steamed custard?

Sugar heated without liquid will, at a temperature of 160° C. (320° F.), melt and form a brown mass, known as caramel, which has a slightly bitter, but agreeable flavor. Caramel dissolved in water, milk, or other liquid is much used as a flavoring for confections and as a coloring matter for gravies and sauces.

Methods of Making Doughs and Batters Light

Doughs and batters are made light by one or other of four methods; viz.:

(1) The introduction of air.

(2) The addition of volatile substances to the dough.

(3) The addition of substances to the dough which will dissociate and thereby yield a volatile gas.

(4) The production of a volatile gas as the result of fermentation.

In all four methods the essential facts depended upon are (1) Air and gas are expanded by heat; (2) that gluten is sticky and expansive and thus serves to hold the air or gas until the heat sets the dough in this porous state. The gas obtained from baking powder and other combinations of acids and alkalies and by the action of yeast is carbon dioxid.

Method 1. Air is generally introduced into doughs, etc., by one of the following ways: (*a*) By beating and rolling the dough as when making paste;

in this process the dough is alternately rolled or pounded and folded over so as to enclose air. The butter, or other fat, by keeping the particles of flour apart, helps to obtain this result. (*b*) The addition of well-beaten eggs to batters. This is a very common method of introducing air, for especially the white of egg, when beaten alone, will take in and hold for a time a large amount of air. (c) Snow is sometimes used as a means of getting air into batters, because the snow crystals contain considerable air.

When air is introduced into batters and doughs in the ways just mentioned, the ingredients must be mixed quickly and the substance cooked at once, for the air will not be held very long as is shown by the rapidity with which the beaten white of egg loses its frothy condition on standing.

Method 2. (*a*) The most common example of the use of this method is the addition of carbon dioxid gas to dough. This is the way in which the aërated breads, much used in England, are made light. (*b*) The addition of an alcoholic liquor, as wine, to a batter helps to make it light, because the heat of the oven volatilizes and expands the alcohol.

Method 3. The substances generally used to obtain carbon dioxid in dough in Method 3 are sodium bicarbonate (baking soda) and an acid. The acid is sometimes obtained by the use of molasses, which contains some free acid and acid salts, or sour milk, which contains lactic acid formed by the fermentation of its lactose. When the lactic acid unites with the soda, the following reaction occurs:

$$\begin{array}{cccccc} \begin{array}{c}\textit{Sodium}\\ \textit{Bicarbonate}\end{array} & & \textit{Lactic Acid} & & \textit{Sodium Lactate} & \\ NaHCO_3 & + & C_2H_5OCOOH & = & C_3H_5OCOONa & + H_2O + CO_2 \end{array}$$

Baking powders are the form in which soda and acid are most commonly used. These consist of sodium carbonate and acid and starch, the starch being used to prevent interaction of the other two constituents before the powder is wanted for use, should it become moist.

The acids commonly used in baking powders are tartaric acid in the form of cream of tartar (potassium bitartrate), phosphoric acid, and alum.

Experiment 51. Object: To show the action of baking powders.

Procedure: (1) Mix ½ teaspoonful of sodium bicarbonate and 1 teaspoonful of cream of tartar in a small beaker. In a second beaker, mix 1 teaspoonful of sodium bicarbonate with 30 c. c. of cold water and stir. In a third beaker, mix 2 teaspoonfuls of cream of tartar and stir. Combine mixtures 2 and 3. What happens? Heat the mixture.

Explain the experiment before reading the following paragraphs.

The effervescence that occurs in Experiment 51 is due to the liberation of carbon dioxid gas; bearing this in mind and that water also is formed, state what salt will be left in a batter when baking powder containing sodium bicarbonate and cream of tartar is used. What salt will be left when phosphoric acid is used?

As shown by the first part of this experiment, no reaction takes place between the alkali and the acid so long as these substances are dry, but as soon as water is added the two combine as described in Chapter XIV. For this reason, baking powders will deteriorate if they are not kept in a dry place.

Another lesson to be learned from this experiment is that chemical action soon ceases and, though the presence of the various solid ingredients of the batter prevent the action occurring and ceasing as quickly as in the experiment, unless the batter is cooked as soon as mixed, the gas will escape before the heat has done its work and the cake, etc., will not be light and porous.

Fate of substances formed in the reaction.—As may be inferred from the preceding paragraphs, there are three things formed in the chemical action between the alkali and the acids used for the purpose under discussion; viz., CO_2, H_2O, and a salt. The CO_2 and the H_2O pass off during the heating, but the salt remains in the food.

Action of salts in the body.—The Rochelle salts, formed when a baking powder contains cream of tartar and the sodium phosphate occurring when the powder contains phosphoric acid, are laxative, but, as they are present in only very small amounts, they have little or no effect on the system. The alum salts that remain when any alum preparation is used are astringent, and it is thought that even in small amounts alum may aggravate any tendency to constipation if people eating food containing it are so troubled.

Experiment 52. A test to detect the presence of alum in baking powder: Dissolve about a teaspoonful of the powder in water, add to this an equal amount of dilute (one-third the usual strength) tincture of logwood, and then add an equal volume of ammonium carbonate solution. If alum is present, the liquid will become a dark blue or lavender color.

Method 4. Yeast is the substance most commonly used for the fourth method of making doughs, etc., light.

Nature of yeast.—Yeasts are microscopic, one-celled plants that multiply by budding (*i. e.*, a small bud or swelling appears on a cell which gradually increases in size and finally splits off). There is more than one variety.

How obtained.—Yeasts are very prevalent in the air and, in the olden days, dough was exposed to the air until enough yeast entered it to cause fermentation. Also some of the fermented dough of one day's baking was often kept to act as a "starter" or "leaven" for the next baking. One great objection to these modes of getting yeast was that other micro-organisms and yeasts with undesirable characteristics entered the dough and frequently interfered with the action of the true yeast and, in this and other ways, spoilt the dough. For this reason, *cultivated yeasts* are now much used; *e. g.*, *compressed* and *dried yeasts*.

Compressed yeast.—This is made by sowing pure yeasts of a desired variety in vats containing material upon which they will thrive. The yeasts grow quickly and in a few hours collect in a scum on the top of the material. This is removed and pressed into cakes. This is one of the best varieties of yeast when used soon after it is prepared, but the cells die from lack of nourishment.

Dried yeast cakes consist of cultivated yeast mixed with other ingredients, chiefly starch, pressed into shape and dried at a low temperature. The drying may kill some of the cells, but, if the cakes are kept in a dry, cool place, the majority of them will survive for a long time. Warmth and dampness will cause temporary activity of the yeast and consequent souring of the cakes.

The brewers' yeast, often used by bakers, consists of the *scum* which grows on the top of the matter undergoing fermentation in the making of beer, etc.

Food required by yeast.—In order that they may grow well, yeasts must be given the kind of food that they require. The principal thing that all the common species require is sugar. This must be in solution and not too concentrated, as is shown by the fact that fruit cooked in a large amount of sugar, as in the making of jam, keeps better than that stewed in a small quantity of sugar.

Action of yeast and cause of this action.—The yeast plants secrete at least three enzymes, one of which facilitates the change of starch to glucose thus: $C_6H_{10}O_5+H_2O=C_6H_{12}O_6$; another causes the fermentation of the glucose and thus changes it to alcohol and carbon dioxid, thus: $C_6H_{12}O_6=2C_3H_5\text{-}OH+2CO_2$, while the third facilitates the oxidation of the alcohol which is thereby changed to carbon dioxid and water. (Give the equation for this reaction.)

It will be seen that in Method 4, as in Method 3, carbon dioxid is the agent which causes the dough to expand.

The addition to dough of sugar and of milk (on account of its lactose) hastens the fermentation, for the sugar provides nourishment for the yeast and makes it more active. For this reason, dried yeast will act more quickly if soaked in sugar solution before being used. The presence of potatoes and rye flour in dough also seems to hasten fermentation; the reason for this is thought to be that their soluble proteins stimulate enzymic action.

The temperature at which dough is kept is of importance. Yeast is most active between 70° F. and 90° F. and it is destroyed at a temperature of 131° F. If dough is allowed to rise too long, acids will develop from the alcohol (acetic acid), from the butter (butyric acid), and from the milk (lactic acid), and thus the dough will be sour. The same thing will happen if the dough is kept at a temperature exceeding 90° F., because the yeast cannot then thrive well and acid-forming species of bacteria which can do so will increase in number.

Experiment 53. Object: To determine the comparative fermentating power of yeasts.

Procedure: Make a dilute mixture of flour and water. Fill the long stems of three fermentation tubes with the mixture. To one add about $\frac{1}{16}$ of a compressed yeast cake dissolved in lukewarm water; to another add a piece of equal size of a dried yeast cake likewise dissolved, and to the third tube add a little brewers' yeast. Mix the contents of each tube by shaking the tubes, see that all the mixture is in the long stem, plug the tubes, stand them in a warm place, and leave them for twenty-four hours. Inspect them

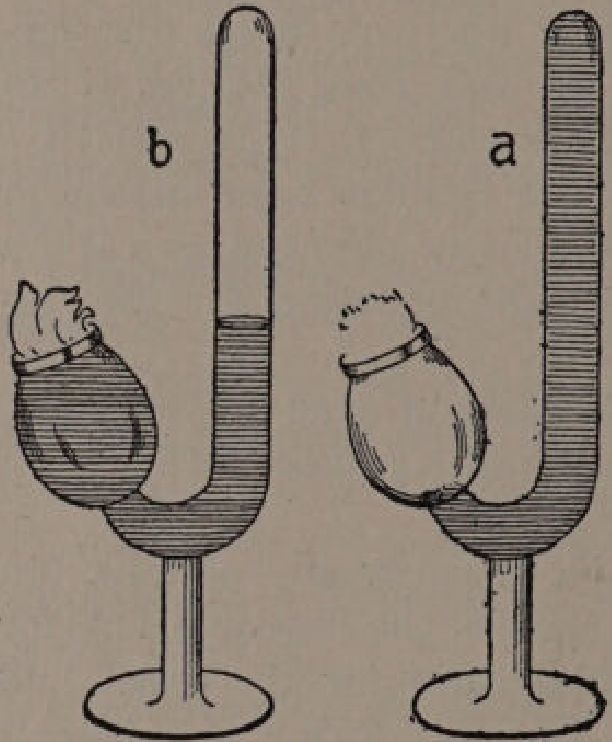

FIG. 65.

(*a*) Liquid is in the long arm of the tube. (*b*) Liquid has been forced into the short arm by the gas.

from time to time. As gas is formed, it rises to the top of the liquid and forces the latter down into the short arm of a tube. Which yeast acted the quickest?

Experiment 54. Object: To see why sugar is added to dough in bread-making.

Procedure: Make a flour-and-water mixture as for the preceding experiment, divide it into two portions; to one portion add about a teaspoonful of cane sugar, to both add $\frac{1}{8}$ of a cake of compressed yeast dissolved in lukewarm water. Put these mixtures into fermentation tubes as in the preceding experiment. In which does fermentation first occur?

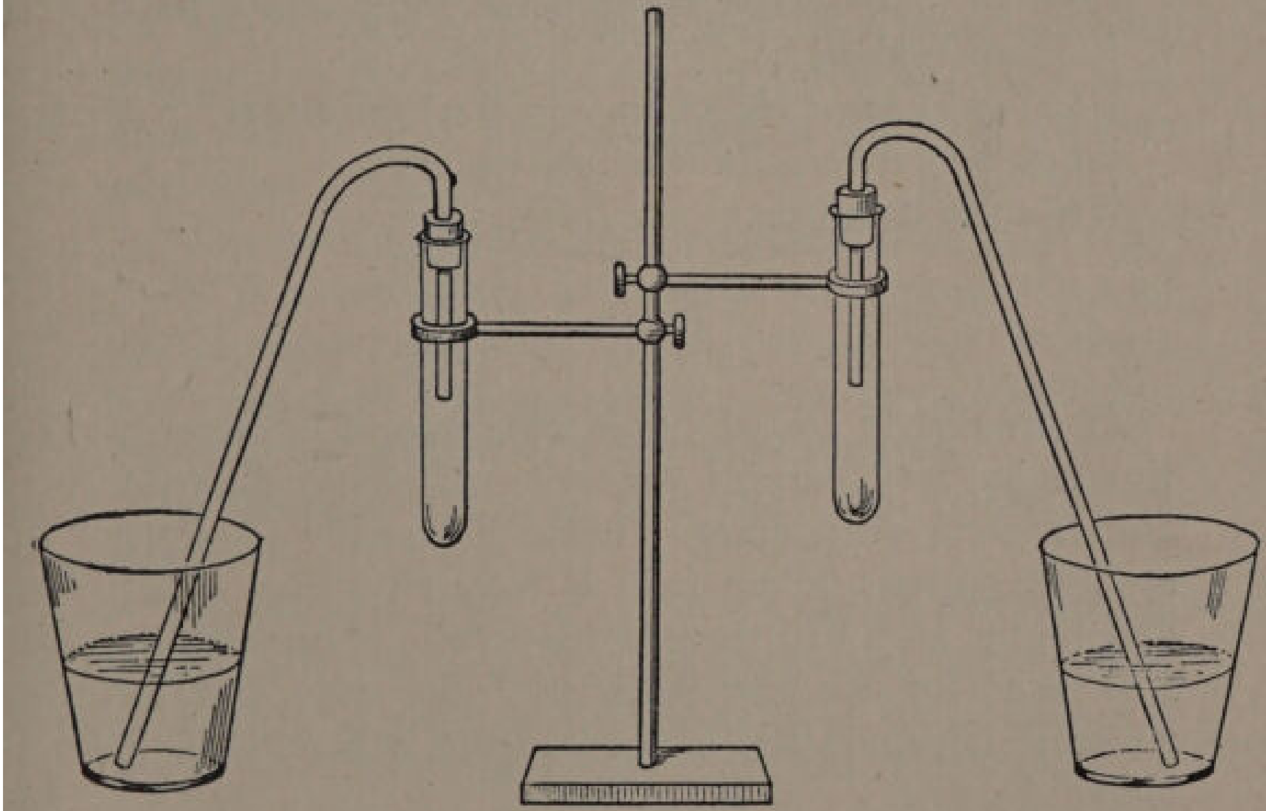

FIG. 66. APPARATUS ARRANGED FOR EXPERIMENT 55.

Experiment 55. Object: To prove that it is the same gas (CO_2) that is obtained by the use of baking powder and yeast.

Articles required: 2 beakers, a test tube with a one-hole cork to fit; glass tubing bent as in Fig. 66; lime water, starch paste, yeast, baking powder, water.

Procedure: Put some of the starch paste, plus ⅛ of an yeast cake dissolved in lukewarm water, into a test tube and lime water into the beaker. Arrange the apparatus as in Fig. 66 and put it in a warm place. Notice the condition of the lime water after twenty-four hours. Put some baking powder in another test tube, some lime water in a beaker, and connect the apparatus as in Fig. 66. Then add some water to the baking powder.

Does the lime water assume the same appearance in the two beakers?

What gas causes the change in the lime water?

What does it form in the lime water?

Write the equation.

Pectin

As previously stated (page 291) pectin is present, though in very varying amounts, in nearly all fruit and in some vegetables. There is usually a larger amount present before fruit is fully ripe than later, and better jelly will therefore be made if part of the fruit used for the purpose is not quite ripe, but, as ripening improves the flavor of fruit, a portion of it should be ripe.

Pectin can be usually extracted from fruit by soaking the latter in warm water, but sometimes it is necessary to boil it (as shown in Experiment 56), and this will always result in the extraction of a larger quantity of pectin. If the cooking is continued too long, however, the pectin will be decomposed to reducing sugars and its gelatinizing property thus impaired.

Pectin, in the presence of vegetable acids, is precipitated or thrown out of solution by such substances as alcohol, salts, and sugar. When a sufficiently large amount of pectin and vegetable acid is present in a liquid the whole solution will be formed into a thick mass known as *jelly* by heating it with sugar. If more sugar is used than there is pectin to unite with, the mass will be soft and gummy or else the excess sugar will crystallize.

The amount of acid present is also of importance for the formation of a firm jelly, and when making jellies from fruit that is not very acid the juice of a fruit that is or a small amount of citric or tartaric acid is sometimes added.

Thus, it can be seen that to make a good jelly from fruit, it is necessary to have proper proportions of pectin, sugar, and acid, and to cook the fruit the exact time required. Consequently, the directions of a reliable recipe must be followed accurately.

Experiment 56. A test for the presence of pectin.

The common test for pectin is to add a little (about 5 c. c.) ethyl alcohol to an equal amount of fruit juice or a water extract of the fruit. The formation of a clear gelatinous precipitate indicates the presence of pectin. At the completion of the experiment the quantity of precipitate obtained from each fruit should be compared.

Any fresh fruit can be used for this experiment, but several should be tried, among others lemons, oranges, and berries, and also carrots; raw fruits, both juices and extracts, and similar cooked fruits should be tested; the juice, peel, and white part of oranges and lemons should be tested separately.

A water extract of the fruit can be usually made by soaking the fruit, after squeezing out the juice, in warm water for about half an hour, but the pectin in the peel and white part of oranges and lemons is hard to extract; therefore, separate the two parts and (keeping them separate) grind or cut them into fine strips, cover these with water, heat this to boiling temperature slowly; strain, cool, and add alcohol to about 5 c. c. of the liquid. Repeat this treatment of the white and peel (using the same portions) until a filtrate is obtained that ceases to show the presence of pectin.

Answer the following questions:

How many extracts were made of the white and of the peel of the orange and the lemon before the filtrate ceased to contain pectin?

Did you find more pectin in the juices or in the extracts from the solid portions of the fruits?

What part of oranges and lemons causes marmalade to jell?

Will soaking oranges before cooking improve a marmalade?

In which fruits did you find the largest amounts of pectin?

Can jellies be made from all fruits containing pectin?

Is there the same amount of pectin in raw as in cooked fruits?

CHAPTER XXI

THE SPOILING, PRESERVATION, AND ADULTERATION OF FOOD

Causes for the Spoiling of Food—Nature of the Organisms that Cause the Spoiling—Chemical and Physical Changes that Occur in the Decomposition of Food Substances—Conditions Necessary to Prevent Contamination and to Preserve Food—Some Common Forms of Food Adulteration.

The Spoiling of Food

Causes of the spoiling of food.—Food spoils because it is attacked by the fungi known as molds, by yeast, or by certain species of bacteria.

Molds

Nature.—Molds, of which there are many varieties, are a species of fungi which, when fully developed, consist of a mass of threads, known as *mycelium*, on which are minute sacs, known as the *spore cases* or *sporangia*, that contain spores.

Multiplication.—Molds multiply and spread because, when their spores mature, the sporangia burst and the spores fall upon the food on which the parent mold is growing or are blown about by the wind and settle on other food, and, also, the spores are carried by insects that have lighted upon molded food.

Experiment 67. Object: To study conditions that favor the growth of molds.

Procedure: (1) Cut some pieces of bread into small —about two inches—squares, put several of these together in layers, wrap them in paper, put them into a tin box, and set the box in a warm place.

(2) Prepare some other slices of bread in a similar manner, store them in a cool dry place.

(3) Put some pieces of bread loosely in a tin box that has a few small holes in either side, for the circulation of air, and place the box in a cool, dry place.

(4) Repeat 1 and 3 with stale bread.

(5) Repeat 1 and 3 with fresh cheese.

(6) Repeat 5 with cheese taken from a piece that has started molding.

(7) Place some pieces of fresh bread where it is exposed to the air and where the sun will shine upon it.

(8) Wrap some cereal in paper and place it in a warm, moist place.

(9) Cut some pieces of fruit, including lemons, and place them in saucers or evaporating dishes. Leave them exposed to the air.

(10) Make a little dilute and saturated sugar solutions, boil them both for one minute, and then let them stand exposed to the air in evaporating dishes.

Let these all stand for a week or longer; examine them from time to time. Notice which ones start to mold first; notice the difference in the appearance of the molds.

As the result of your observations answer the following questions:

Is there any apparent change in the appearance of the food? Has it grown less in quantity?

Is there any odor?

Will bread keep better where there is a free circulation of air or where air is excluded?

How do temperature and moisture affect the growth of molds?

Under what conditions must food be kept to prevent it molding?

Will dried food mold as quickly as that containing considerable water?

Does the presence of a molded piece of food influence the spoiling of the remainder of food kept in the same place?

Does fruit acid inhibit the growth of mold?

Why will preserved fruits keep better when a large amount of sugar is used in their cooking?

Examine some of the different molds under the microscope. To do this, remove a small portion of mold, with a platinum wire or point of a penknife, from the food upon which it is growing and place it on a glass slide; try to keep it in the same position as that in which it grew on the food. Examine it with the low-power objective of the microscope.

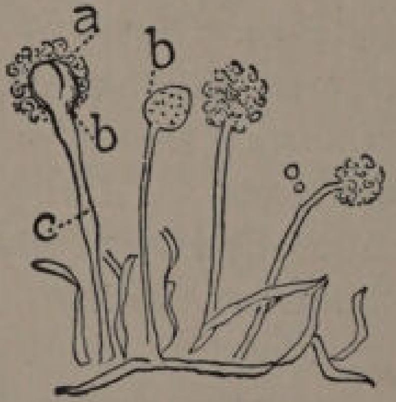

FIG. 67.

SPECIMEN OF MUCOR, A COMMON FORM OF MOLD MAGNIFIED.

(a) Spores.
(b) Sporangium.
(c) Mycelium.

Yeast

The nature and action of yeasts were discussed in Chapter XX.

Bacteria

The bacteria concerned in the spoiling of food are the various species classed as *saprophytes; i. e.*, those which feed on non-living organic matter and in doing so cause it to putrefy, ferment, or decay.

Result of infection of food by molds, yeasts, and bacteria.—These organisms use the food upon which they fall for their nourishment, but they cannot do so while it is in its natural condition any more than human beings can assimilate the food they eat until it has been digested. Therefore, just as happens in the human alimentary canal, these organisms, by virtue of enzymes which they secrete, cause the disintegration of food into simpler substances and from these they remove what they require for their subsistence. Some of the substances formed during the process are gaseous and these soon pass from the material; this is one reason for the loss of bulk and accounts for the odors that develop. Some of the other disintegration products are liquid and others are soluble and go into solution in the water present; thus food undergoing putrefaction, fermentation, etc., is usually softer than unspoiled food. When the molecules of the food material are broken, they no longer contain all the elements that they can hold in combination and they unite with oxygen or other elements; oxidation is thus associated with decomposition processes. Read the following chemical formulæ and decide which of the processes thus portrayed are associated with oxidation:

$$C_6H_{12}O_6 = 2C_3H_6O_3$$
$$C_3H_6O_3 = CO_2 + C_2H_5.OH$$
$$C_2H_5,.OH + O = CH_3.COH + H_2O$$
$$CH_3.COH + O = CH_3.COOH$$
$$CH_3.COOH + 3O = 2CO_2 + H_2O$$
$$\overset{\text{Ammonia Carbamate}}{CO_2N_2H_6} - H_2O = \overset{\text{Urea}}{CON_2H_4}$$
$$C_6H_{12}O_6 = 2H_2 + CO_2 + \overset{\text{Butyric acid}}{C_4H_8O_2}$$

Effect of changes caused by molds, etc., on the wholesomeness of food.—Certain molds and bacteria are purposely cultivated in food in order to produce the desired flavor, *e. g.*, the molds of Stilton and Roquefort cheeses, but in such case special cultures of the organisms are used.

The common consequences of mold infection is the production of a musty disagreeable flavor in the food. As a rule, the substances formed by molds in food are not poisonous, unless the condition proceeds to such an extent that putrefactive bacteria are also at work, as usually occurs in advanced stages of molding. Consequently, when the molding is not extensive, if the mold and the material lying directly below it are removed, the food can be used. If it is to be kept for any length of time, however, it must be boiled for at least twenty minutes in order to destroy any spores that may have fallen into it.

The substances produced by yeast are not unwholesome and fermentation can be arrested in the same way as molding, but if the fermentation process continues the food will become so acid that it will be unpalatable.

The substances produced by bacteria that cause fermentation are similar to those produced by yeast, but food in which putrefactive changes have occurred has, when eaten, caused severe and even fatal poisoning. Just which germs are responsible for this so-called *ptomain poisoning* is not known, but it is known that they thrive better in the absence of air than in its presence, and in warm places. Food which has been in cold storage and later kept in a warm place is particularly apt to be so infected.

Conditions and Food Most Conducive to the Life of Molds, etc.

As shown by Experiment 67, molds grow best in dark, damp places where there is no free circulation of air. Why moving air interferes with the growth of molds is not wholly understood, but it is thought that disturbance of their threads by the wind may interfere with their growth and also that air currents increase the rapidity of the evaporation of moisture from the surface of the food and thus renders it an unfavorable ground for the molds. Thus, molds grow much more rapidly in tightly closed vessels or cupboards than in those through which there is a current of air; in food piled up than in that spread out; in food filled with cavities, like bread and cheese, than in that with a smooth surface. Though dampness favors mold development, very moist food, as milk and meat, are not as likely to be attacked by molds as by bacteria. Molds, however, will grow upon food far too dry to be used by bacteria (*e. g.*, cereals and dried fruit) when the atmosphere surrounding the food is moist.

Yeasts require food containing sugar or starchy substances that they can change to sugar. They grow best at a temperature between 70° and 90° F. Their growth is arrested at a temperature below 40° F. and they are soon killed at boiling temperature.

The majority of saprophytic bacteria thrive best at a temperature of 70° to 90° F., and are killed by moderate heat—140° to 160° F.—if it is continued for from thirty minutes to one hour, but there are some species that thrive best at a temperature of about 140° F. and can be killed at boiling temperature only if it is maintained for at least one hour. The growth and action of all species are inhibited at a temperature below 40° F. Bacteria grow better in darkness than in light—direct sunlight will kill many varieties. Bacteria, even more than yeasts and molds, require moisture for their development; many kinds will not grow in food containing less than 25 to 30 per cent. moisture.

Practical Application of the Knowledgement of the Requirements of Molds, Yeasts, etc., to the Keeping of Food

Since molds and bacteria prefer darkness to light, kitchens and pantries should be very light rooms and food cupboards and storerooms should be provided with ventilators and they should be as light as possible without raising their temperature. All these places and all utensils into which food is put must be dry and clean.[1] Places where food is stored must be cold.

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[1] Such utensils as bread and cake boxes and cereal jars should be frequently scalded and exposed to the air. Milk utensils should be well aired and sterilized after use.

Keeping Fruit and Vegetables

Such vegetables as potatoes and beets have somewhat different requirements from the majority of foodstuffs, for if given light and air, they will germinate and sprout, therefore they are kept in the dark, packed in bins or barrels. Carrots, turnips, and parsnips keep better if buried in sand. For all these vegetables, the storeroom must be cool and dry. Winter squash, unlike the other vegetables, keep best if spread out in a rather warm place.

It is most important for the preservation of fruit and vegetables that they be kept dry while in storage. Though, as shown by Experiment 67, some foods mold more quickly when rolled in paper, because they are then not exposed to air currents, soft, porous paper wrapped around fruit that is packed in boxes seems to prevent it molding and decaying, probably because the paper absorbs the moisture from the surface of the fruit and prevents that in which decay has started coming in contact with sound fruit. Grapes are packed in sawdust for the same reasons. Drying fruit with a soft cloth also helps to keep it fresh. The peel of fruit helps to protect it from molds and bacteria; thick-skinned fruits, like oranges and winter apples, can be, with proper care, kept in their natural state for months, but thin-skinned fruit, like cherries, will not keep in good condition for any length of time. Only fruit in perfect condition should be stored; if the skin is bruised or broken, molds, etc., can easily enter and, as seen in Experiment 67, one piece of molded food will soon infect that which is sound.

Means of Preserving Food

The means usually taken to preserve food are (1) sterilizing it and keeping it under sterile conditions. (2) Keeping it in a temperature below 40° F. (3) Drying it. (4) Adding substances to it that will inhibit the growth of organisms. When used for this purpose such substances are called *preservatives*.

Sterilization.—Food to be preserved in this way must be submitted to a high enough temperature, for a sufficient length of time, to kill bacteria. It has been found that the time and temperature necessary vary with different foods; therefore, when preserving or canning food, the directions of a reliable recipe book regarding these matters should be followed. Some foods, *e. g.*, peas, beans, and corn, require a very long time—from one to two hours—at a temperature of 212° F., to be rendered sterile. Food will be better sterilized if put into cold water and brought to boiling point, than if put at once into boiling water. Sterilized food, if it is to be kept, must be put into sterile jars, while it is still hot, and sealed at once. The sealing must be perfect, so that it will be impossible for germs to enter. If, for instance, hard rubbers are used for jars into which jam is put, the sealing will not be perfect since the tightness with which the cover can be screwed down depends upon the softness of the rubber.

In canning factories, tin jars are usually used. The fruit and vegetables, with or without previous boiling, are put into the cans and the covers are soldered on. There is a small hole in each cover, to allow the air to escape as it becomes expanded by the heat. The cans are placed in the heating apparatus and left the required length of time. At the conclusion of the sterilization, they are removed and the holes are at once covered with solder.

Cold.—In order to preserve food by cold for any length of time the temperature must be kept even and below 40° F.—in fact, it is necessary for the preservation of many foods to keep the temperature as low as 33° F. Even at this temperature, bacterial action, though inhibited, is not entirely prevented and thus food cannot be kept in cold storage indefinitely without deterioration occurring.

Drying.—This is one of the most widely used means of preserving food—meat, fish, fruit, and vegetables all being preserved in this way. Some of the fruits that are frequently thus preserved contain enough sugar to aid in their preservation. Salt is often used, in addition to drying, in the preservation of meat and fish and these foods are sometimes smoked, as well as dried, since the creosote of the wood acts as a preservative by destroying germs upon the surface of the food and by so impregnating it that bacteria cannot grow in it.

Drying is usually accomplished by exposing the food to heat from various sources and in various ways, or else by submitting it to hydraulic pressure and thus removing much of its water.

Drying toughens the fiber of foods, both those of animal and vegetable origin, and causes considerable loss and change of flavor. In order to minimize the result of toughening, nearly all dried foods are soaked in water for several hours before being cooked. This does not, however, apply to the cereals, which are dry by nature and subjected to little if any artificial drying.

Preservatives.—These are often classed as *legitimate preservatives* and *adulterants*.

By legitimate preservatives are meant substances that will inhibit the growth of germs, but in the quantities used have no injurious effect upon man. Those in common use are salt, sugar—used in sufficient quantity to be concentrated—, spices, and vinegar.

By adulterants are meant substances such as borax, benzoic acid, salicylic acid, and formaldehyd. Some people contend that these substances should not be classed as adulterants for reasons that will be seen later.

"According to the law of many of the States of America, a food is declared to be adulterated under the following conditions:

"First, if any substance or substances have been mixed with it, so as to lower or depreciate or injuriously affect its quality, strength, or purity; second, if any inferior or cheaper substance or substances have been substituted wholly or in part for it; third, if any valuable or necessary constituent has been wholly or in part abstracted from it; fourth, if it is an imitation of or is sold under the name of another article; fifth, if it consists wholly or in part of a diseased, decomposed, putrid, infected, tainted, or rotten animal or vegetable substance or article, whether manufactured or not, or, in the case of milk, if it is the product of a diseased animal; sixth, if it is colored, coated, polished, or powdered, whereby damage or impurity is concealed, or if, by any means, it is made to appear better or of greater value than it really is; seventh, if it contains any added substance which is poisonous or injurious to health. Provided: that the provisions of this act shall not apply to mixtures or compounds recognized as ordinary articles or ingredients of food, if each and every package sold or offered for sale bear the name and address of the manufacturer and be distinctly labeled under its distinctive name and in a manner so as plainly and correctly to show that it is a mixture or compound and is not in violation with definitions fourth and seventh of this section."[1]

Those who contend that the antiseptic preservatives mentioned on page 375 should not be classed as adulterants base their assertion on the fact that the preservatives are often of greater commercial value than the foods in which they are used and that in the small quantities in which they are used they are not injurious to health. Those who take the opposite view contend that even if the preservatives do themselves cost more than the food, they allow the manufacturers to use cheaper—even partly spoiled—food, for preserving, canning, etc. Also they hold that it has never been proved that even small quantities of salicylic acid, borax, etc., if constantly used, are not to some extent injurious, especially to children and invalids.

A few simple experiments to detect the presence of chemical preservatives in food[2]:

The chemical reagents required for these experiments are as follows:

Turmeric paper
Alum solution, 10%
Iron alum (crystal or powder form)
Sulphuric acid
Hydrochloric acid
Tincture of iodin
Potassium permanganate 1%
Ethyl alcohol
Chloroform
Boric acid or borax
Ammonia water

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[1] Bulletin 100, U. S. Department of Agriculture.

[2] The majority of these experiments have been taken from Bulletin 100, U. S. Dept. of Chemistry: "Some Forms of Food Adulteration and Simple Methods for their Detection."

As in the tests to discover the constituents composing food, the majority of tests to discover food adulterants consist in the use of reagents that will, if the matter sought is present, combine with it and form a substance that will be precipitated or that will give a special color reaction. As a rule, the tests are better performed with the material being tested in solution or, at least, in a semi-liquid condition. Solid and semi-solid foods are therefore usually dissolved, and, if necessary, macerated in distilled water before the reagent is added.

Salicylic Acid

Salicylic acid is used for preserving condiments, sauces, and fruit products of all kinds—jellies, marmalades, etc.,—and it is a common constituent of the various powders sold as *preservatives.*

Experiment 68. Put about 60 c. c. of the liquid to be tested into a small flask, add a few drops of sulphuric acid, and shake the flask for two or three minutes. Filter this solution into another small flask. Add about 30 c. c. of chloroform and mix the two liquids by a somewhat vigorous rotary motion (shaking is to be avoided as it produces an emulsion which is difficult to break up). Pour the solution into a beaker. Let it stand until the chloroform settles at the bottom and then, with a medicine dropper or pipet, transfer as much of the chloroform as possible to a test tube. As salicylic acid is soluble in chloroform, it will now, if there was any present in the food, be in the chloroform. To the chloroform, add an equal amount of water and a small piece (a little larger than a pinhead) of iron alum. Shake the mixture and then allow it to stand until the chloroform again settles at the bottom. If salicylic acid is present, a purple color will develop at the top of the liquid.

Benzoic Acid

Benzoic acid is often used for preserving jams, jellies, tomato catsup, and similar articles.

Experiment 69. Extract the sample with chloroform according to the directions given for salicylic acid, but, after removing the chloroform, put it into a beaker and put this in a warm place or a water bath. Let it stand thus until the chloroform evaporates. If benzoic acid is present, it will be seen in the form of characteristic flat crystals on the bottom of the beaker and these, if heated, will emit the unmistakable odor of benzoic acid.

Boric Acid and Borax

Boric acid—known also as *boracic acid*—and its sodium compound borax are often used to preserve animal foods, such as sausages, butter, and milk.

Experiment 70. To prepare solid food for this test, macerate it with distilled water, strain it through fine white cotton cloth, and then filter the liquid thus obtained through filter paper.

To prepare butter, place a heaping teaspoon in a beaker, add about 10 c. c. of boiling, distilled water, stand the beaker in a water bath until the butter is thoroughly melted. Stir this with a teaspoon and place the beaker, with the spoon in it, in a cold place until the butter becomes solid. Remove the spoon and the butter (which adheres to it) and filter the liquid through paper.

To prepare milk, pour about 60 c. c. into a small flask with about twice that amount of a 10 per cent. solution of alum, shake the mixture vigorously, and then filter it.

Test: Take about 5 c. c. of the liquid obtained as described, add and mix thoroughly 5 drops of concentrated hydrochloric acid. Dip a strip of turmeric paper into the liquid and then hold the paper in a warm place until it dries. If there is boric acid or borax in the liquid, the paper, when dry, will be a bright cherry red and a drop of ammonia will change the red color to dark green or greenish black. If too much hydrochloric acid is used, the turmeric paper may become a brownish red, even though there is no boric acid in the solution; in such case, however, ammonia changes the color to brown, just as it does turmeric paper which has not been dipped into the acid solution.

Formaldehyd

Formaldehyd, being volatile, is useless as a preservative for food that is to be kept for any length of time, but it is often used in warm weather to prevent milk souring. It is probably the most commonly used milk preservative at the present time, though boric and salicylic acids are also used.

Experiment 71. Pour about 30 c. c. of milk into a flask and an equal quantity of concentrated hydrochloric acid. Add a small piece of ferric alum (about the size of a pinhead) and mix the fluids by rotating the flask between the hands. Then place the flask in a water bath containing boiling water and let it stand for five minutes (do not keep heat under the bath during this time). If formaldehyd is present, the mixture will become purple.

Copper Sulphate

Copper sulphate is sometimes added to canned green vegetables, especially peas, to improve their color.

Experiment 72. To about a teaspoonful of the mashed vegetable add three teaspoonfuls of water and 30 drops of strong hydrochloric acid. Stir the mixture with a piece of wood and set the beaker in a water bath. Drop a bright iron wire nail (tin carpet tacks will not do) into the mixture and keep the water surrounding the beaker boiling for twenty minutes. Stir the mixture frequently. At the end of twenty minutes, remove the nail and examine it. If there was any appreciable amount of copper in the mixture, the nail will be coated with it.

Other Forms of Adulteration

The mixing of a cheap article with a more expensive one is an even more common form of adulteration than the use of chemicals for preservation. For example, cheap cereals and flours are mixed with more expensive ones; chicory, ground beans, and the like are mixed with coffee; starch and saccharin are added to jellies; cheap fruits, as apples, are mixed with rarer ones in jellies and jams. As apples are as wholesome as any other fruit, their addition to jams, etc., is not considered an adulteration if their presence is stated on the label.

Experiment 73. To detect starch in jellies, jams, etc. In order to detect starch in any substance, the iodin test is used, but, so that change of color may be perceived if it occurs, it is necessary to decolorize colored substances such as jellies. To do this, dissolve the material, heat it to boiling point, and add potassium permanganate, drop by drop, until the liquid is colorless. Then cool the mixture and add iodin. The presence of starch is indicated by the iodin becoming blue.

Adulteration of Cereals and Flours

Unchanged starch grains have the characteristic structure of the seed in which they were formed, and as the seeds of different plants vary in shape, microscopic examination of starch, etc., is a means often employed to detect adulteration of one grain with another, as, for instance, wheat flour with corn flour, which is cheaper.

Adulterations of Coffee

The presence of adulterants in coffee can be often detected by examining a sample under a magnifying glass. When seen under the glass, the coffee particles appear quite different from the coffee substitutes employed. Chicory, especially, having a dark gummy appearance, stands out in strong contrast to the coffee. All the usual coffee substitutes, except chicory, contain starch and, therefore, the iodin test is another means of detecting adulterations other than chicory. A weak infusion must be used so as not to obscure the color change. Still another method of testing ground coffee is to put a small amount of coffee into a bottle half-full of water, shake the bottle vigorously, and then let it stand quietly. Pure coffee contains a large amount of oil and, consequently, it will float while the substitutes sink to the bottom of the flask.

Adulterations of Butter

The so-called *spoon test* is often used to distinguish pure butter from renovated or process butter and oleomargarin.

Experiment 74. Take a small piece of the butter sample in a large spoon and hold it over the Bunsen flame. If the sample is fresh butter, it will boil quietly with the evolution of a large amount of froth. Oleomargarin and process butter, on the other hand, crackle and sputter as they melt and boil.

The Adulteration of Salad Oil

The adulteration of salad oil with cotton-seed oil is not uncommon.

The Halophen test is very often used to detect this adulteration.

Experiment 75. Procedure: Put about 5 c. c. of oil into a test tube and add an equal quantity of both amyl alcohol and carbon disulphid. Cork the tube and place it in a beaker of hot water. Leave it thus for half an hour; keep the water hot during this time, but do so by the addition of boiling water when necessary; do not bring the tube near a flame, for carbon disulphid is excessively inflammable. At the end of that time, if even a small percentage of cotton-seed oil be present, the mixture will be of a distinct reddish color, and if the sample consists largely or entirely of cotton-seed oil, the color will be deep red.

The Adulteration of Milk

The more common methods of adulterating milk are: The use of preservatives, the addition of water, the removal of a part of the cream.

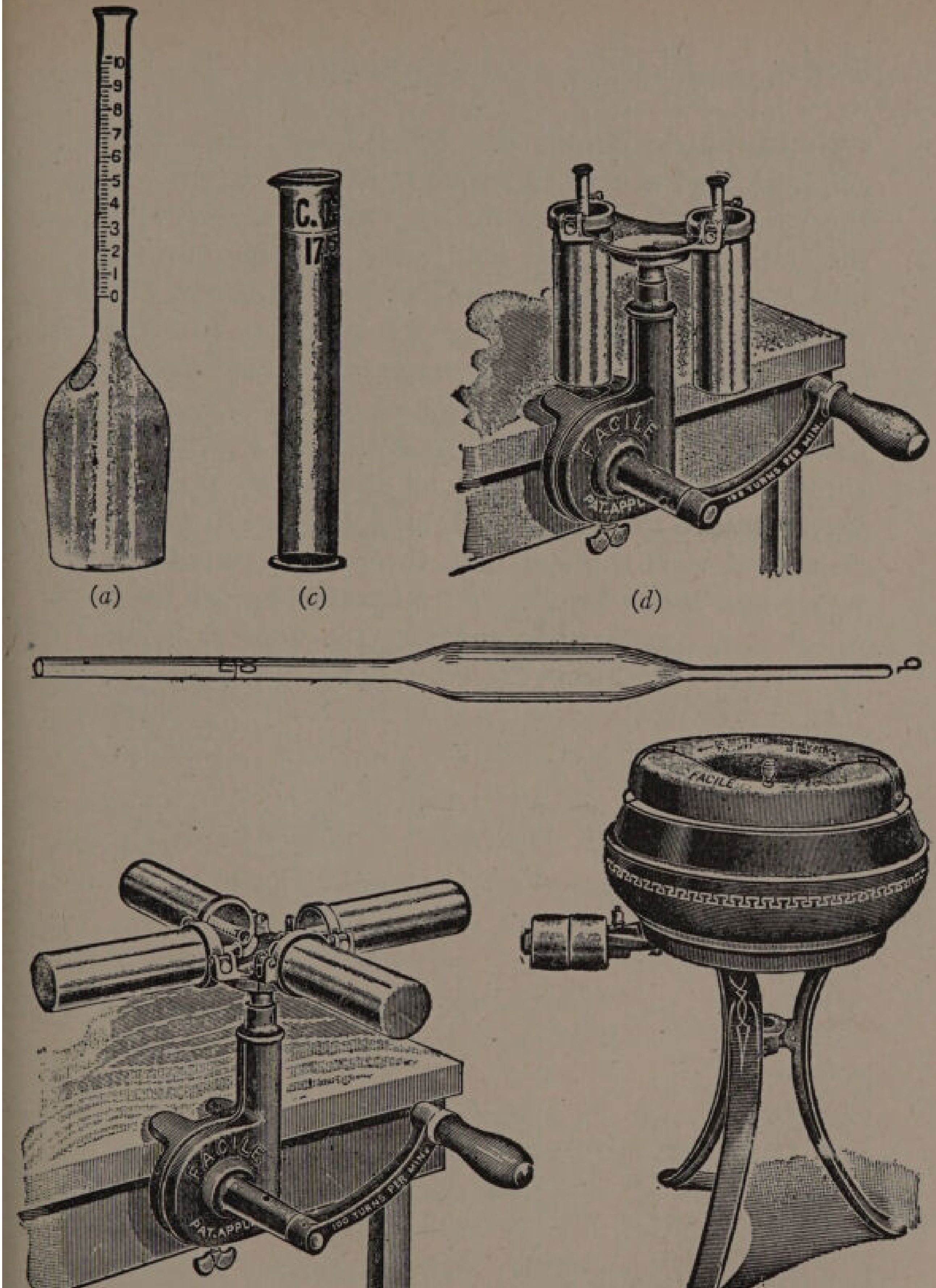
The preservatives most commonly used are formaldehyd and boric acid. Means of detecting these have been already discussed. The presence of preservatives in milk is particularly objectionable because (1) it is so much used for invalids and young children who may be injured by such preservatives; (2) the action of lactic acid bacteria is easily inhibited and the souring of milk thus retarded, but germs which cause putrefaction are not so easily affected and the changes which they cause in milk, while infinitely worse, are not so easily detected. Consequently, milk containing preservatives may also contain injurious products of putrefaction. It is such substances in milk that are often responsible for the gastro-intestinal disturbances in children that are so common in summer.

The presence of water in milk is determined by the use of a lactometer or hydrometer. The specific gravity of water, it will be remembered, is said to be 1 and that of ordinary whole milk, as compared with water, is 1.027 to 1.033. If the milk contains more than 4 per cent. cream it may be slightly lower than these figures, but much difference indicates the presence of the liquid with the lower specific gravity, *i. e.*, water.

The Babcock test.—This test is very commonly used for estimating the per cent. of fat in milk. The test depends upon the fact that the organic matter of milk, except fat, is decomposed by sulphuric acid, while, practically, the only change in the fat is that it is set free.

The articles required for the test are: A milk test bottle (see Fig. 66; this bottle has a long neck, graduated in such a way that the numbers 1, 2, 3, etc., represent 1, 2, 3, etc., per cent. fat and the small lines between the numerals represent $\frac{2}{10}$ of 1 per cent.); a graduated pipet; an acid measure; a centrifugal machine; the milk to be tested; hot water; sulphuric acid with a specific gravity of 1.82.

Procedure: The milk to be tested is measured with the pipet. To do this, hold the pipet between the thumb and middle finger, put the pointed end of the pipet into the milk and the other end in the mouth, and, by suction, draw the milk up above the marks on the stem. Then, quickly, put the index finger over the upper end of the pipet, to keep the milk from running out. Next, hold the pipet vertically, with the 17.6 mark on a level with the eye and, by slightly relaxing the pressure of the finger covering the upper end of the stem, allow the milk to flow out until it is at the 17.6 mark. Then place the pointed end of the pipet in the mouth of the milk bottle, holding both bottle and pipet obliquely, remove the finger from the upper opening and let the milk run down the neck of the bottle. If the milk is put in otherwise, the air in the bottle may form bubbles in the milk and cause it to overflow. If the milk does not all leave the pipet, blow into the upper end of the tube. Fill the acid measure to the 17.5 mark with sulphuric acid and add this to the milk slowly, holding the bottle



(a) (c) (d)

(e) (f)

FIG. 68. APPARATUS FOR BABCOCK MILK TEST.
(a) Milk Bottle. (b) Pipet. (c) Acid Measure.
(d), (e), (f) Different varieties of Centrifugal Machines.
(d) A two-bottle Hand Machine (not in action).
(e) A four-bottle Hand Machine (in action).
(f) An Electric Machine.

obliquely while doing so. (N. B., remember that sulphuric acid is very corrosive.) Whirl the bottle in the centrifugal machine for five minutes. Remove the bottle, add enough hot water to bring the fat into the narrow neck. The fat can be easily discerned, for it is lighter in color than the rest of the liquid, which holds the decomposed organic matter in solution. After the addition of the water, replace the bottle in the centrifugal machine and whirl it for three minutes. This brings all the fat to the top. Then remove the bottle and, holding it with the fat on a level with the eye, note the marks opposite its upper and lower levels. To estimate the per cent. of fat it is necessary to subtract the number below the level of the fat from the top reading. For example, if the upper level of the fat is on a line with the mark 5.2 and the lower level with the mark 1. 2, the correct reading will be 4 per cent. since $5.2-1.2=4$. This would show that the milk tested contained 4 per cent. fat.

CHAPTER XXII

CHEMISTRY OF DIGESTION

The Nature of the Changes Occurring in the Digestion of the Various Food Stuffs—The Organs in Which These Changes Occur—The Factors and Conditions Influencing Digestion—Nature of Enzymes, Zymogens, and Kinases.

Reason for digestion.—With the exception of the monosaccharids and the disaccharids, organic food substances are unable to pass through animal membranes and, consequently, they are of no use to the body until they undergo certain changes which transform them into substances that will go into solution and pass through the walls of the blood-vessels and lacteals. For it is only after food has passed into the blood and been carried by it to the tissues that it benefits the body. The various changes that food undergoes in this process of preparation are classed as *digestion.*

Mineral matter, monosaccharids, and disaccharids are readily soluble in water and the two first mentioned are quickly absorbed from the stomach and the intestines. Disaccharids will be absorbed unchanged if eaten in large quantities, but, when this happens, they are quickly eliminated from the system in the urine, the tissues being unable to utilize them.

Nature of digestion.—The digestion of proteins and carbohydrates consists of a process of hydrolysis—*i. e.*, decomposition due to the absorption of, and chemical combination with, water. The digestion of fats consists of a preliminary splitting of the fats into fatty acids and glycerin, followed by the saponification of the fatty acids.

Results of hydrolysis and saponification.—As the result of hydrolysis, complex molecules are divided and simpler substances thus formed; for example, one molecule of a disaccharid ($C_{12}H_{22}O_{11}$) by hydrolysis gives 2 molecules of a monosaccharid $2(C_6H_{12}O_6)$. There is only one splitting of the molecules in the hydrolysis of the disaccharids, but the polysaccharids $(C_6H_{10}O_5)_x$[1] pass through several stages before they are completely hydrolyzed. The substances yielded at the different stages in the digestion of starch are dextrins (known, according to the colors which they give with iodin, as erythrodextrin 1, 2, or 3 and achrodextrin, see page 398), maltose, and glucose. The protein molecule also undergoes several splittings, thereby giving rise to such substances as metaproteins, proteoses, peptones, and amino acids. The nature and the results of saponification whereby fats are digested have been already discussed (Chapter XVI.).

How hydrolysis and saponification can be brought about.—Starches, sucroses, and proteins can be hydrolyzed and thereby made to undergo the changes described in the preceding paragraph by boiling them with acid and by the use of ferments. The ferments used to digest starch outside the body are usually obtained from plants; one of the best known is diastase of malt, a substance that is produced during the ger-

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[1] The x, it will be remembered, signifies that the molecules of polysaccharids contain an unknown number of molecules with the construction demonstrated by the formula.

mination of certain seeds and that is contained in malt. The ferments used to digest proteins are extracted from the stomach or pancreas of animals, especially pigs. Within the body, hydrolysis is caused by the ferments or enzymes contained in the digestive juices. Saponification, outside the body, is brought about by boiling fats with alkalies and, within the body, by an enzyme contained in the pancreatic juice and the alkaline salts of the pancreatic and intestinal juices and the bile.

Nature of the digestive juices.—The digestive juices secreted by the salivary glands, by the gastric and intestinal glands, and the pancreas consist of water, salts, mucin, and specific enzymes. The reaction of the saliva is, normally, neutral or faintly alkaline; that of the gastric juice, acid (due to the presence of hydrochloric acid); and that of the pancreatic and intestinal juices, alkaline. The bile, which is secreted from the blood by the liver, consists of water, bile salts, and inorganic salts, mucin, cholesterin, lecithin, fat, and bile pigments.

The bile pigments are derived from hemoglobin; it is thought that when red blood-corpuscles become disintegrated the freed hemoglobin is brought to the liver, where, under the influence of the liver cells, its iron is split off and it is converted into bilirubin or biliverdin, the characteristic pigments or coloring matter of bile. Normally, bile has an alkaline reaction.

Enzymes, Zymogens, Kinases, Hormones

The nature and action of these substances which are so essential to digestion and metabolism are as yet very imperfectly understood.

Enzymes.—Enzymes are protein-containing substances, of unknown composition, produced by living cells, both animal and vegetable. Enzymes act as catalyzers—*i. e.*, they hasten chemical reactions, but they do not themselves enter into the reaction. Nearly all the processes, both of building and of disintegration that take place in animals and plants are controlled to some extent by enzymes. Enzymes are easily influenced by temperature; those of animal origin act best at body temperature; those obtained from plants usually work better at higher temperatures than this, but all enzymes are destroyed by extreme heat and are rendered more or less inactive by cold. The majority of enzymes are also very sensitive to the reaction of the medium in which they are contained; some needing an acid and others an alkaline or neutral medium. They are all destroyed by strong acids and alkalies. Nearly all enzymes are specific in their action—*i. e.*, they each have their special work to perform and they do only that work.

According to their action, the majority of enzymes found in the animal body have been classified as follows:

(1) Hydrolytic enzymes—those which cause hydrolysis. To this class belong the proteolytic or protein-splitting enzymes, the lipolytic or fat-splitting enzymes, the amylolytic or starch-splitting enzymes, and the invertases or sugar-splitting enzymes.

(2) The coagulating enzymes—such as the thrombin of the blood and the rennin of the gastric juice.

(3) The oxidases—enzymes in the body tissues that are essential for oxidation processes.

(4) Reduction enzymes—those which remove oxygen from matter.

Zymogens.—Certain enzymes, notably the rennin of the gastric juice and the trypsin of the pancreatic juice, are secreted in an inactive form. This is necessary because the pepsin and trypsin, being formed for the purpose of hastening the digestion of proteins, would, if present in the gland secreting them in the active state, cause the digestion of the gland. Such inactive enzymes or precursor of enzymes are called *zymogens*. The zymogen of pepsin is called *pepsinogen* and that of trypsin, *trypsinogen*.

Kinases.—Enzymes, or other substances, that produce the necessary change in zymogens to render them active, are called *kinases*—for example, the enzyme enterokinase of the intestinal juice which activates the trypsinogen of the pancreatic juice.

Hormones, secretins.—Hormones are chemic substances that are produced in certain organs and then absorbed and carried by the blood to other organs which they stimulate. Of this nature is the secretin, a chemic substance of unknown composition, which is produced from a substance secreted in the intestinal mucous membrane, called *prosecretin*, by the action upon it of the acid material entering the intestine from the stomach in the course of digestion. This secretin, when absorbed by the blood, is carried to the pancreas, liver, and possibly to the intestinal glands, and produces or increases the activity of these glands. A secretin is also produced in the mucous membrane of the pyloric portion of the stomach, which, being absorbed and carried by the blood to the secretin cells of the gastric glands, excites them to activity.

The names of the digestive juices, the enzymes they contain, the action of the latter upon food, and

stomach, the gastric juice is less acid than it is in the central portion. In these cells, it is thought, the NaCl is by some means broken into Na and Cl and the Cl combines with H, thus forming HCl. The per cent. of HCl in the gastric juice during digestion is usually estimated at from 0.3 to 0.5 per cent. A much lower or higher degree of acidity interferes with digestion.

The principal purposes of the hydrochloric acid are as follows: It activates the pepsinogen and assists in the digestion of proteins; it regulates the passage of food from the stomach; it acts as a kinase in that it acts upon the pepsinogen and changes the prosecretin contained in the mucous membrane of the intestinal wall to secretin; it acts as a disinfectant. HCl arrests the action of ptyalin when the food with which that enzyme became mixed in the mouth becomes saturated with the acid. However, as there is little or no acid in the fundus of the stomach, the ptyalin may continue its action on starch while the latter remains in that part of the organ and, as the mechanical action, which churns and pushes the food onward, is not very great in that region, some of the starch may remain there for twenty minutes or more.

How hydrochloric acid regulates the passage of food from the stomach.—The reaction of secretions in the stomach, when the latter is empty, is, normally, neutral or slightly alkaline, the HCl being secreted only after food enters the stomach. As protein and HCl have a strong affinity for each other, the protein of the food and the acid formed at once combine and it is only after all the protein has become saturated with the acid that there is any free HCl in the gastric juice. As soon as this happens, under the influence of the acid, the pyloric sphincter muscle, which encircles the orifice between the stomach and the duodenum and is usually in a state of contraction, relaxes and some of the acid chyme[1] passes into the intestine. This makes the alkaline contents of the intestine acid and the effect of acid on this side of the pyloric sphincter is to cause it to contract. The pylorus remains closed until the intestinal contents are alkaline when it once more relaxes and the procedure is repeated.

Factors influencing the secretion of digestive juices.—Various experiments have shown that the secretion of the digestive juices is very largely controlled by (1) the nervous system and (2) secretins. The stimulation of the parts of the nervous system that are concerned with the digestive process is brought about in several ways; *e. g.*, the smell and taste of food—provided that they are agreeable—and the movements of the jaw in mastication all stimulate various nerve-endings in and around the mouth. This results in the stimulation of certain nerve-centers in the medulla oblongata from which impulses pass to the secretory cells of the salivary and gastric glands and stimulate them to secrete. The entrance of food into the stomach also excites nerve-impulses which stimulate the secretion of gastric juice, but the principal sources of gastric stimulation are thought to be the secretins formed in the pyloric mucous membrane and, during digestion, from food. The secretin formed, as already explained, in the small intestine after the acid chyme enters it from the stomach, stimulates the pancreas and probably the

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[1] This is the name given to the substance resulting from gastric digestion.

intestinal glands to secrete and it increases the activity of the liver.

The entrance of food into the small intestine also causes the sphincter muscle closing the duodenal opening of the common bile duct to relax, thus allowing the passage of bile from the liver and gall bladder into the intestine.

The kinds of food eaten influence the nature and quantity of the digestive juices; fats, for instance, retard gastric secretion, and sugars, condiments, meat extracts, and alcohol increase it. The amount of the various enzymes present in the digestive juices varies with the kinds of food eaten, and it is thought probable that there are specific stimuli contained in food or produced during digestion whose action is of a kind to arouse reflexly the secretion best adapted for the digestion of the kind of food that is eaten.

Digestion is aided also by the contractions of the stomach and intestinal walls. Weakness of the muscle tissues of these organs, or any form of interference with their mechanical action, is likely to be followed by digestive disturbances and, frequently, by an accumulation of fermentation or putrefactive products. These are due to bacterial action upon food that remains too long within the stomach and intestines.

Action of bacteria in digestion.—After birth, there are always a large number of different kinds of bacteria present in the alimentary canal. The varieties that exist in the stomach and intestines are, under normal conditions, unless present in abnormally large numbers, not only not injurious to man, but of use, for they help to protect the body against the invasion of harmful varieties, and they help in decomposing food into simpler substances, causing, in fact, much the same kind of changes as are produced by enzymes. Under abnormal conditions, however, as when the mechanical action of the stomach and intestines are defective and food remains too long in these organs, or when, for any reason, the digestive power is impaired, fermentation and putrefactive changes may be carried too far and gases produced, thus causing flatulence. Also various substances may be formed that, when absorbed into the blood, are more or less injurious to the system. Of these substances, the best known is indol, which is produced in the intestine by the excessive putrefaction of proteins. It is changed, it is thought, in the liver, to indican and as such is excreted by the kidneys. The amount of indican in the urine is therefore an index of the degree of putrefaction occurring in the intestines.

Kinds of bacteria in the alimentary canal.—The principal varieties of bacteria present in the alimentary tract are those which promote fermentation and putrefaction. The various organisms that produce these two processes do not thrive well together and, normally, bacteria which produce fermentation are found in the stomach—these requiring considerable air—and those which produce putrefaction and thrive best without air, in the intestine. So great is the inhibitory action of these organisms on each other that lactobacilli—organisms that produce lactic acid fermentation—are used medicinally to lessen intestinal putrefaction.

Products of digestion.—As the result of digestion, there are found in the intestine monosaccharids, soaps, peptones, and amino acids. Also, there will be present varying amounts of undigested and imperfectly digested food, mineral matter, and water.

The products of digestion, mineral matter, and much of the water are principally absorbed from the small intestine. The changes that occur in the soaps and protein derivatives during absorption will be discussed in Chapter XXIII. The sugars and proteins pass directly into the blood, via the capillaries that are contained in the villi covering the inner wall of the intestine. These capillaries connect with the portal vein through which these absorbed substances pass to the liver. A small portion of the fat may enter the circulation in this way, but by far the greater part only reaches the blood after it has passed through the lymphatic vessels that extend from the lacteals (the small lymph vessels in the villi of the intestine) to the thoracic duct which opens into the left innominate vein between the left subclavian and the left internal jugular veins.

Feces.—All the matter not absorbed passes slowly through the intestine and is eliminated from the rectum. Chemical and microscopical examinations of feces show its principal constituents to be food residue mixed with digestive juices, dead and living bacteria, and a small amount of waste matter formed in metabolism. The last passes from the blood as it flows through the intestinal blood-vessels.

Experiments on Digestion

Nature of the experiments.—Many of the experiments used in studying the progress of digestion are the same as those described in Chapter XVIII. The iodin and reduction tests (see pages 289 and 290) are used in studying the digestion of starch, since if, when iodin is added to the solution being tested, a blue color develops, it shows that starch is still present and that digestion has not started; the development of a purple color shows that the starch has been changed to erythrodextrin No. 1, which is the first stage in the digestion of starch; a red color indicates the presence of erythrodextrin No. 2, a red-brown that of erythrodextrin No. 3, and a loss of color from the iodin shows that the starch has been changed to either achrodextrin, maltose, or glucose; which of the three can be determined by finding the amount of solution that it takes to reduce 10 c. c. or other given quantity of Fehling's solution since it takes 0.5 gm. of glucose, 0.8 gm. of maltose, and about 1.0 gm. of dextrin to reduce 10 c. c. of Fehling's solution. Consequently, there will be a greater amount of precipitate when the solution tested contains glucose than when it is maltose or dextrin that is present, and a sufficiently accurate judgment for class purposes can be made by comparing the relative color and amount of precipitate formed in all the tests of an experiment. Of course, the same amount of the solution to be tested and of the reducing solution must be used in all tests that are to be compared.

The reduction tests are used also to see if disaccharids have been hydrolyzed, since they, like starch, are changed to monosaccharids in digestion.

The biuret and xanthoproteic tests—described on pages 283 and 284—are much used in studying the digestion of proteins for they give different color reactions with undigested and digested proteins. The biuret reagent assumes a violet color in the presence of undigested proteins and a rose color in a solution of peptones, while the solutions used for the anthroproteic test become orange-colored in solutions containing undigested or only slightly digested proteins, but this color does not develop after the proteins have been changed to proteose and peptones.

The progress of the digestion of fats can be observed by the use of litmus, for, as soon as the fats in a solution are split to fatty acids and glycerin, blue litmus will become red, as it always does in the presence of acids.

N. B.—The measurements for these tests must be especially accurate, very slight inaccuracies being sufficient to spoil many of them. Before beginning experiments number the tubes as described on page 11. Differences in color in many of the tests is very slight and can be usually detected only by comparison, therefore all results must be kept until a test is completed.

Experiments 76 to 78.

Object: To study the digestion of starch.

Articles required: Test tubes, test-tube racks, Bunsen burner, iron stand, c. c. measures, funnels, filter paper, ice, starch paste made as directed on page 21, distilled water, iodin, Fehling's solution, saliva. Before attempting to collect the saliva, rinse the mouth with water. Thinking of something good to eat, chewing a little paraffin or moving the jaws as though doing so, or inhaling a little ether vapor will increase the flow of saliva. Collect the saliva in a test tube, add an equal amount of distilled water to it and filter it.

Experiment 76. Object: To test the action of saliva on starch.

Procedure: Into a tube marked 1, pour 10 c. c. of starch paste, into tube 2 pour 5 c. c. of starch paste and 5 c. c. of saliva. In tube 3 put about 5 c. c of crackers mashed in a small amount of distilled water and 5 c. c. of saliva. Shake the tubes containing the saliva so as to mix the latter thoroughly with the starch or cracker. Put all three tubes into a water bath and keep them at a temperature of 40° C. After five minutes pour 2 c. c. from each of these tubes into separate tubes, divide each 2 c. c. and test one portion with iodin and one with Fehling's solution. Repeat this procedure every five minutes for thirty minutes. Keep all the tests in order until the completion of the experiment and then compare the colors.

Why did the crackers show the presence of glucose before the starch?

Experiment 77. Object: To test the effect of acids and alkalies on the ptyalin.

Procedure: Into each of three tubes put about 1 c. c. of starch paste and 1 c. c. of saliva. To tube 2 add 2 c. c. of 0.4% HCl and to tube 3, 2 c. c. of 2% Na_2CO_3. Put the three tubes into a water bath and keep them at a temperature of 40° C. for twenty minutes, then divide the contents of each tube into two equal portions and test one portion with iodin and the other with Fehling's solution.

Experiment 78. Object: To test the effect of temperature on the enzymes of the saliva.

Procedure: Put about 1 c. c. of starch paste into each of three tubes. To tube 1 add an equal amount of saliva, to tube 2 add an equal amount of saliva that has been boiled for at least three minutes, put both these tubes into a water bath and keep them at a temperature of 40° C. for twenty minutes. To tube 3 add saliva that has been chilled by keeping the tube containing it between pieces of ice. After mixing the starch and saliva, surround the tube with ice, leave it thus for twenty minutes. At the end of twenty minutes divide the contents of the three tubes into equal portions and test one portion of each with iodin and the other with Fehling's solution.

Experiments 79 to 80.

Object: To test the digestion of proteins.

Articles required: The same apparatus, but not the solutions, as for the experiments with starch. Also, concentrated nitric acid, ammonia, sodium hydroxid 10%, copper sulphate solution 1%, hydrochloric acid 0.5% and 2%, sodium carbonate solution 10%, pepsin solution, (see page 21), fibrin. Dried fibrin, derived from blood, can be obtained at dealers in scientific apparatus. Meat or egg can be used instead of fibrin, but the results are not as quickly obtained. The biuret test cannot be used with egg, since egg albumin sometimes gives, unlike other undigested proteins, a rose color.

Experiment 79. Object: To test the effect of hydrochloric acid and alkali on pepsin.

Procedure: Mark tubes 1, 2, 3, 4. Into each tube put a small piece of fibrin and about 4 c. c. of neutral pepsin solution. To tube 2, add an equal amount of 0.5% hydrochloric acid, to No. 3 add an equal amount of 2% hydrochloric acid,[1] and to tube 4 an equal amount of 10% sodium hydroxid solution. Let the four tubes stand in a water bath kept at 40 ° C. for twenty minutes. Watch the changes that occur in the fibrin during this time. At the end of twenty

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[1] Be careful not to use too much acid, for though, as this test should show, strong acid interferes with the action of pepsin, it will, if stronger than necessary, itself aid in the hydrolysis of the protein.

minutes pour off about 2 c. c. of the fluid from each tube into separate tubes and apply either the biuret or xanthoproteic tests (see pages 283 and 284). Let the fibrin remain in the remaining 1 c. c. of solution until that in tube 2 is dissolved—this may require several hours—and then filter the solutions into separate tubes, divide each one into equal parts, and test one portion with the biuret and the other with the xanthoproteic test.

Account for the results.

Experiment 80. Object: To study the effect of temperature on pepsin.

Procedure: Mark tubes 1, 2, 3. Put a small piece of fibrin into each tube. Add enough HCl to 10 c. c. of neutral pepsin solution to make it a 0.2% acid solution. Divide this into three parts. Surround the tube containing one portion with ice and leave it thus until it is thoroughly chilled, then pour it over the fibrin in tube 3 and put this tube between ice. Boil another portion for at least three minutes and then pour it over the fibrin in tube 2, add the third portion to the fibrin in tube 1. Put tubes 1 and 2 into a water bath and keep them in a temperature of 40° C. for twenty minutes. Then filter the contents of all three tubes into separate tubes and apply the biuret test.

Experiments 81 to 83. Object: To study the action of the pancreatic juice.

Articles required: The same apparatus used in the experiments with the saliva and gastric juice, also starch solution, pancreatic solution—see page 21—hydrochloric acid 2%, sodium carbonate solution 2%, sodium hydroxid solution 10%, litmus powder, milk.

Experiment 81. Object: To test the effect of pancreatic solution on starch.

Procedure: Repeat Experiments 76 to 78, using pancreatic juice instead of saliva.

Experiment 82. Object: To test the effect of pancreatic solution of protein.

Procedure: Repeat the Experiments with pepsin solution, using the pancreatic solution instead of the pepsin solution, and in Experiment 79 substitute sodium carbonate solution 2 per cent. for the 0.5% hydrochloric acid solution.

Experiment 83. Object: To test the action of pancreatic juice on fat.

Procedure: Into each of 2 test tubes put 10 c. c. of whole milk and enough litmus powder to color the milk a decided, but not dark, blue. To tube 2 add 2 c. c. of pancreatic juice. Put both tubes into a water bath and keep them at 40° C. After a short time, it will be noticed that the blue color of the milk in tube 2 changes to pink. Why does it do so?

CHAPTER XXIII

THE CHEMISTRY OF ABSORPTION AND METABOLISM

Changes that Occur in Food Substances during Absorption—Nature of Metabolism—Composition of the Blood—Changes that Occur in Food Substances during Metabolism—Some Causes and Results of Defective Metabolism—The Fuel Value of Foods—Food Requirements.

Absorption

The changes that occur in food substances during absorption.—Of the substances formed during digestion from disaccharids, polysaccharids, proteins, and fats, only monosaccharids are usually found in the general circulation. Therefore, it is thought that during absorption and, probably, in the liver chemical changes the opposite of those occurring in digestion take place. These changes result in the transformation of the soap into fats and of peptones and amino acids into blood proteins—the serum albumin or globulin. The passage of these substances into the blood is thought to be partly due to osmotic pressure, but both their passage into the blood and their rebuilding are accomplished by the activity of cells in the mucous membrane of the intestine. Whether these cells form enzymes, or whether the enzymes with which the food became mixed in the process of digestion by a reverse[1] action aid in this synthesis, are still debated questions.

Metabolism

Every living organism may be likened to a chemical laboratory, whether it be plant or animal, whether it be large and composed of millions of cells or so small that it cannot be seen without the aid of a microscope; for in the cells of all such organisms chemical reactions are constantly taking place whereby the tissues are built and repaired and, in the case of animals, heat and energy are generated to keep their body machinery in motion. The classification of such processes in plants was mentioned in Chapter XVIII.; those occurring in animals are spoken of as *metabolism.* Some of the reactions occurring in metabolism can be easily performed in any chemical laboratory, but after a great many years of study, even the most skillful chemists and physiologists have but an indefinite knowledge of the nature of some of the processes.

Catabolism and anabolism.—The chemical reactions of metabolism, such as hydrolysis and oxidation, which result in the disintegration of complex substances into simpler compounds, are classed as *catabolism* and the changes are said to be *catabolic* (from the Gr. *kata—down* and *metabole=change*), and the synthetic changes in which simple substances are put together to form complex compounds are classed as *anabolism* and the changes are said to be *anabolic* (Gr. *ana=up* and *metabole=change*).

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[1] A common characteristic of enzymes is that an accumulation of the products of their action will check their activity and even cause them to reverse the nature of their action.

Composition of the blood.—Since the blood is the reservoir for the chemical supplies that the body uses for the numerous chemical reactions constituting metabolism, it may be well to recall its composition before going further with the study of the reactions.

The constituents of the blood are usually classified as follows:

- Water
- Solids
  - Corpuscles
    - White
    - Red
    - Blood-platelets
  - Proteins
    - Fibrinogen
    - Serum globulin or paraglobulin
    - Serum albumin
  - Extractives[1]
    - Nitrogenous substances such as urea, uric acid, creatin
    - Fats
    - Lecithin
    - Cholesterin
    - Glycogen
    - Glucose
  - Salts
- Gases
  - Oxygen
  - Carbon dioxid
  - Nitrogen

Also, there are in the blood ferments, enzymes, internal secretions, antitoxins, and similar substances.

The composition of the blood in the portal vein differs at times from that in the other vessels because

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[1] Substances, other than proteins, that may be extracted from dried residue of blood by water, alcohol, or ether.

it contains all the substances absorbed from the intestine and some of these the liver removes or changes before they enter the general circulation. For instance, glucose, over the amount necessary to provide the blood with a content of about 0.1 or 0.15 per cent., is removed by the liver cells and changed to glycogen, which is stored chiefly in the liver, but to some extent in the muscles.[1] Also, putrefactive products, such as indol, that are absorbed from the intestine are changed to substances less likely to injure the system before they enter the general circulation.

Derivation of the blood constituents.—The constituents of the blood, other than the corpuscles, antitoxins, and similar substances, represent: (1) food matter absorbed from the intestines; (2) oxygen taken from the lungs;[2] and (3) matter, useful and otherwise, that has been absorbed from the tissues and glands.

Fate of the food constituents of the blood.—There is a constant osmosis of substances from the blood, through the capillaries into the tissues, and this matter is either used by the cells for their nutrition or building or else it unites with oxygen and is broken down, just as fuel is when it combines with oxygen in a furnace. The result of this oxidation is the same as that which occurs in the furnace, viz., the production of heat. In the body, some of this heat is used for energy to keep the body machinery—heart, lungs, etc.—in action as well as to provide the body with power to perform its external work.

[1] After the death of an animal the glycogen and glucose in the muscles and blood are soon oxidized.

[2] This is contained principally in loose combination with the hemoglobin of the red blood-corpuscles.

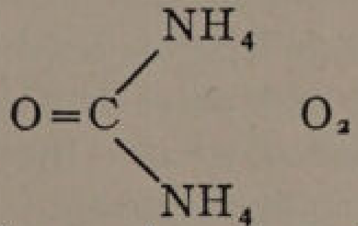
Intermediate Stages in the Metabolism of Food Products

Glucose, as already stated, is changed to glycogen and as such stored in the liver, but in health the blood maintains a constant content of about 0.1 to 0.15 per cent., and as it parts with its sugar to the tissues the glycogen is reconverted to glucose. In the tissues, the glucose ($C_6H_{12}O_6$), it is thought, is split to lactic acid ($C_3H_6O_3$). and then undergoes a series of oxidations with the formation first of various acids and finally of carbon dioxid and water which are eliminated through the lungs. If more food is eaten than the body requires for its heat and energy supply, glucose can be changed to fat and stored as fatty tissue.

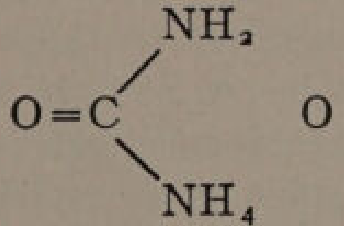
When metabolism of glucose is defective, as in diabetes, much of the glucose does not undergo complete oxidation and there is an accumulation of glucose and of acids in the body, and glucose is eliminated in the urine.

Fat may be stored as part of the body fat; it may be synthesized with other substances to form more complex body constituents—such as lecithin—or, as is in any case the final fate of fat, it is oxydized to carbon dioxid and water and during the process yields heat and energy. The stages through which fat passes before it is completely oxidized are but imperfectly understood. When its metabolism is defective, diacetic acid and acetone are formed.

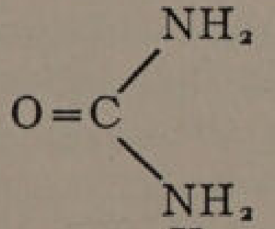
The **proteins** that are eaten as food vary considerably in their composition and many of them are very different from body tissue, but, since they are all composed of different combinations of amino acids and the combinations are parted in the process of digestion and in metabolism, the body is provided with the same fundamental constituents as those which compose its own substance and which its cells are capable of putting together, not in the order in which they were originally, but as they are needed for the muscle tissue and other proteins of the body. Since the proteins are the only foods containing these amino acids, they are the only ones that can be used in the body for the repair and building of its protein constituents. The molecules of protein not used for these purposes are split to amino acids and these in turn are split to form ammonia (NH_3) and organic acid radicals. The ammonia, it is thought, unites with carbonic acid (CO_2+H_2O), which is being constantly formed in the course of metabolism, forming ammonia carbonate. This, by a loss of one molecule of water, leaves ammonia carbonate, which in turn, by a loss of water, yields urea. This is excreted by the kidneys.



Ammonia carbonate



Ammonia carbamate



Urea

The organic acid radicals left after the removal of the ammonia group consist, like fats and carbohydrates of, COH and, like those substances, they may be oxidized to carbon dioxid and water, yielding heat and energy, or they may be synthesized to fat or glycogen.

The nucleo-proteins—which are the characteristic proteins of cell nuclei, both those of food and of the human body—when metabolized, give rise to substances that are known as *purin bases*. These, on further decomposition, yield uric acid. It is thought that uric acid is formed both in the liver and throughout the tissues of the body. As the acid is formed, most of it unites with alkaline substances, forming salts—*urates*—which are excreted in the urine.

The formation of salts in the body.—Though acids are being constantly formed in the body in the course of metabolism, the body fluids, except the gastric juice and the urine, are alkaline or neutral. This is because the acids are either oxidized and thus broken down to substances such as carbon dioxid and water, which can be eliminated, or else they are neutralized by uniting with alkaline substances obtained from food and carried through the body by the blood. Many of the salts excreted in the urine, sweat, and feces are formed within the body in this way.

Factors influencing metabolism.—That the many chemical reactions which occur in metabolism can do so at body temperature is due to the presence in the tissues of various enzymes, which are manufactured by the body cells, and internal secretions carried thither by the blood from the glands in which they are secreted.

Results of defective metabolism.—As shown in the preceding paragraphs, broken-down body tissue and also food substances that are not built into tissue go through several stages of disintegration before they are eliminated. The intermediate products, only a few of which were mentioned, are principally acids and nitrogen compounds. If, for any reason, metabolism is defective, the final stages of disintegration are likely to be delayed and abnormal compounds are often formed, thus there is likely to be an accumulation of more or less injurious substances within the body.

Some causes of defective metabolism.—Diseases, such as anemia, and conditions which prevent the body getting its necessary amount of oxygen, will result in defective metabolism, since many of the disintegration processes are due to oxidation. Conditions of the tissues which retard enzyme formation, or diseases of those ductless glands which are concerned with metabolism, will interfere with it, as will also disturbances of certain parts of the nervous system, since metabolism and heat regulation, like all other vital processes, are, both directly and indirectly, controlled by this system.

Heat Value of Food and Food Requirements

The study of these things belongs to dietetics, but a brief survey of them here may be of advantage.

How the heat value of a food is determined.—The heat value of food is calculated and recorded in the same manner as that of other substances as described in Chapter III. It was determined by burning foods in a form of apparatus known as a *bomb calorimeter*. By this means it was found that each gram of carbohydrate eaten should yield to the body 4 calories; each gram of fat, 9 calories; each gram of protein, 4

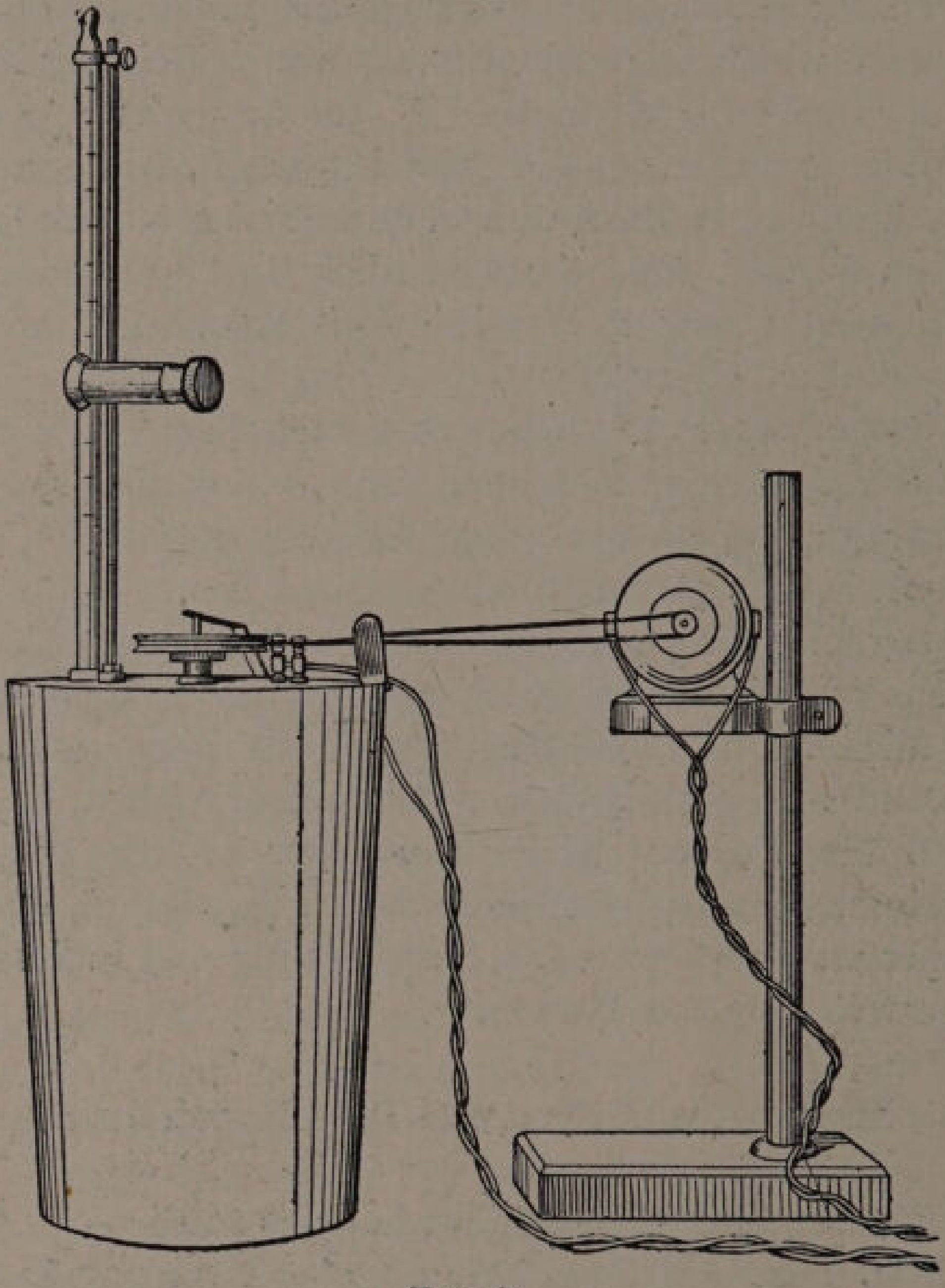


FIG. 69.

A VARIETY OF CALORIMETER USED FOR ASCERTAINING THE HEAT VALUE OF FOODS.

calories. If proteins were as fully oxidized in the body as they are in the calorimeter, they would have a higher food value, but it will be remembered that

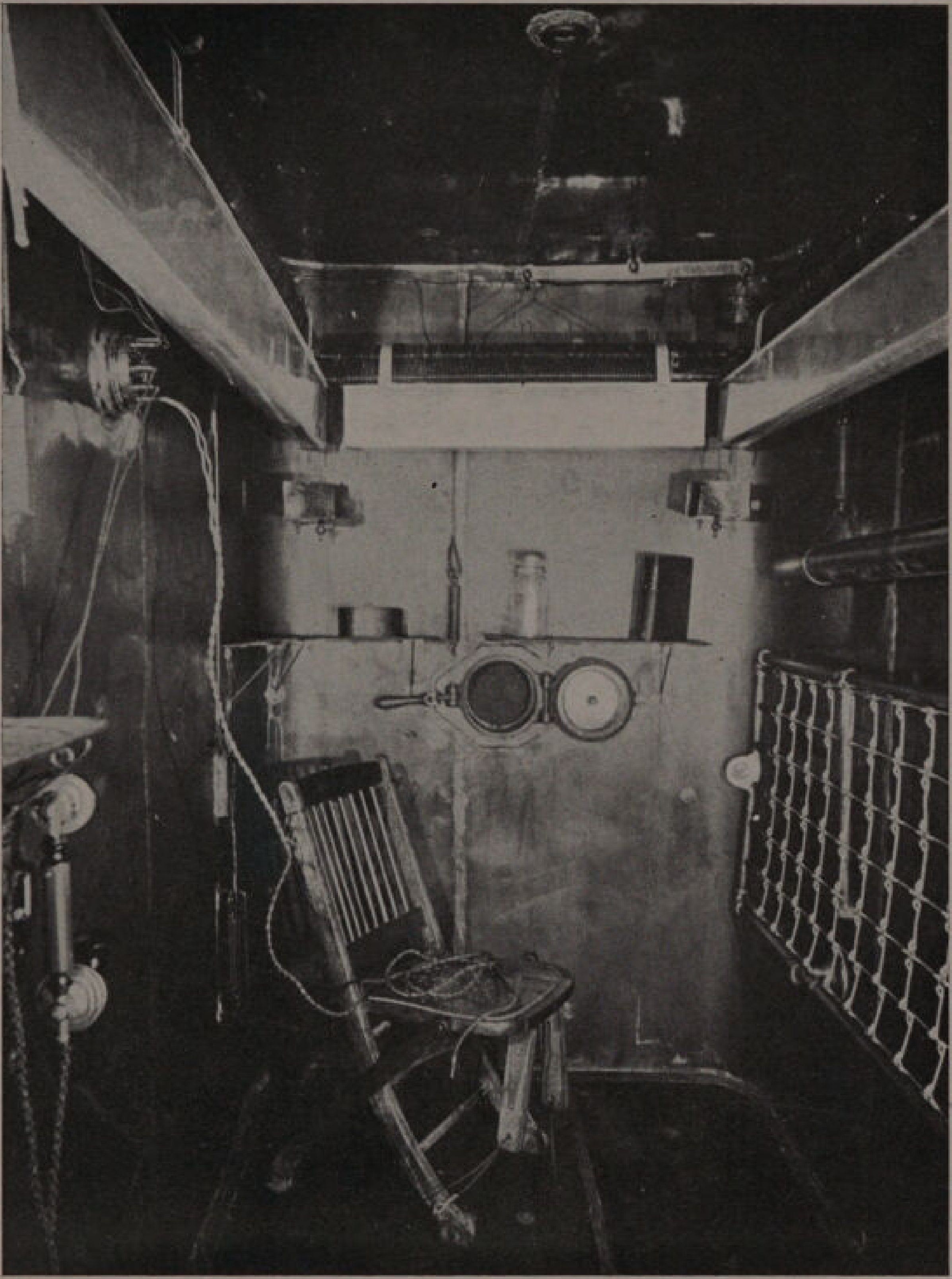


FIG. 68.—THE INTERIOR OF THE RESPIRATION CHAMBER. VIEW TAKEN FROM THE WINDOW.

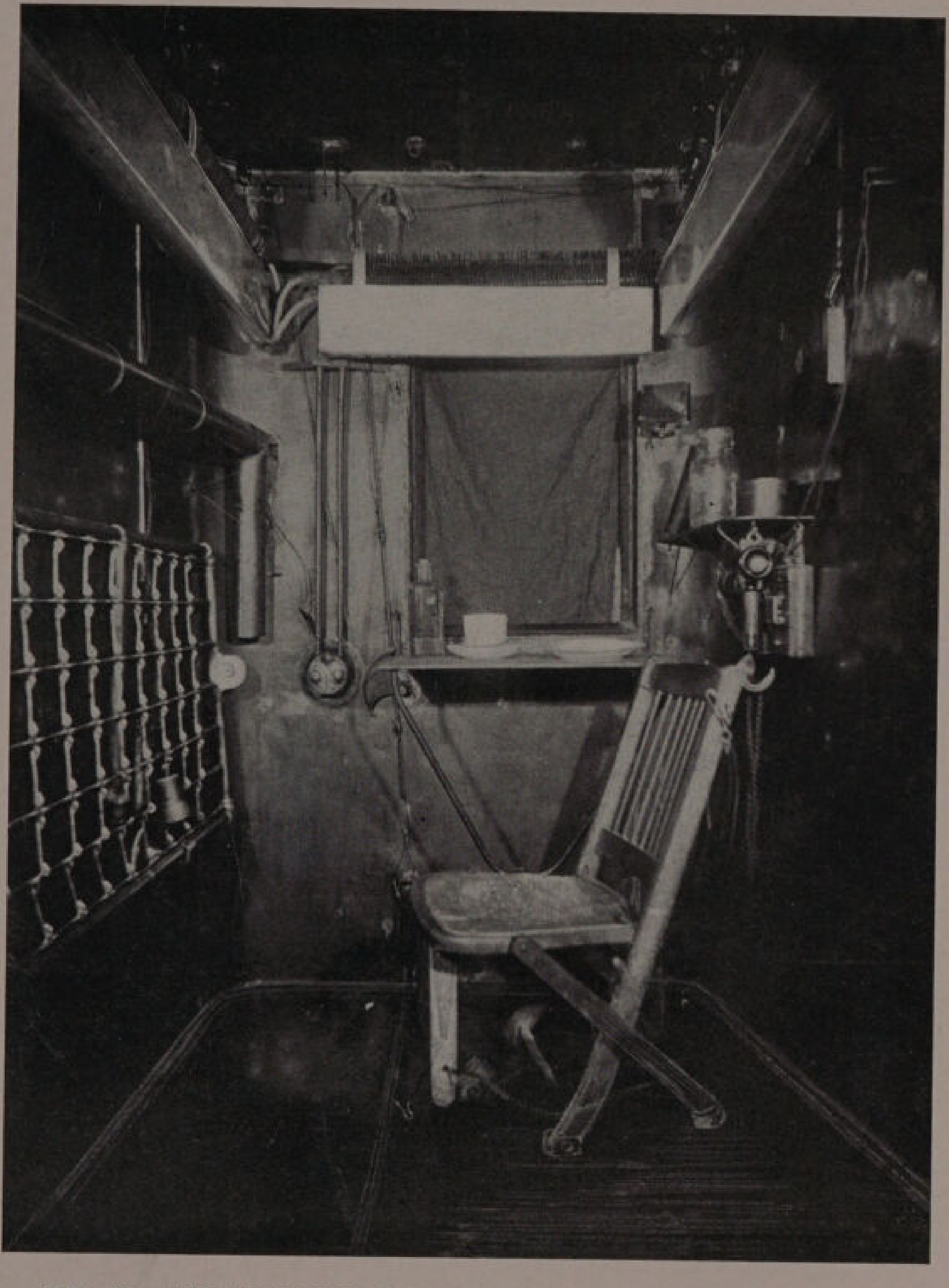


FIG. 69.—THE INTERIOR OF THE RESPIRATION CHAMBER. VIEW TAKEN FROM THE REAR OF THE CHAMBER.

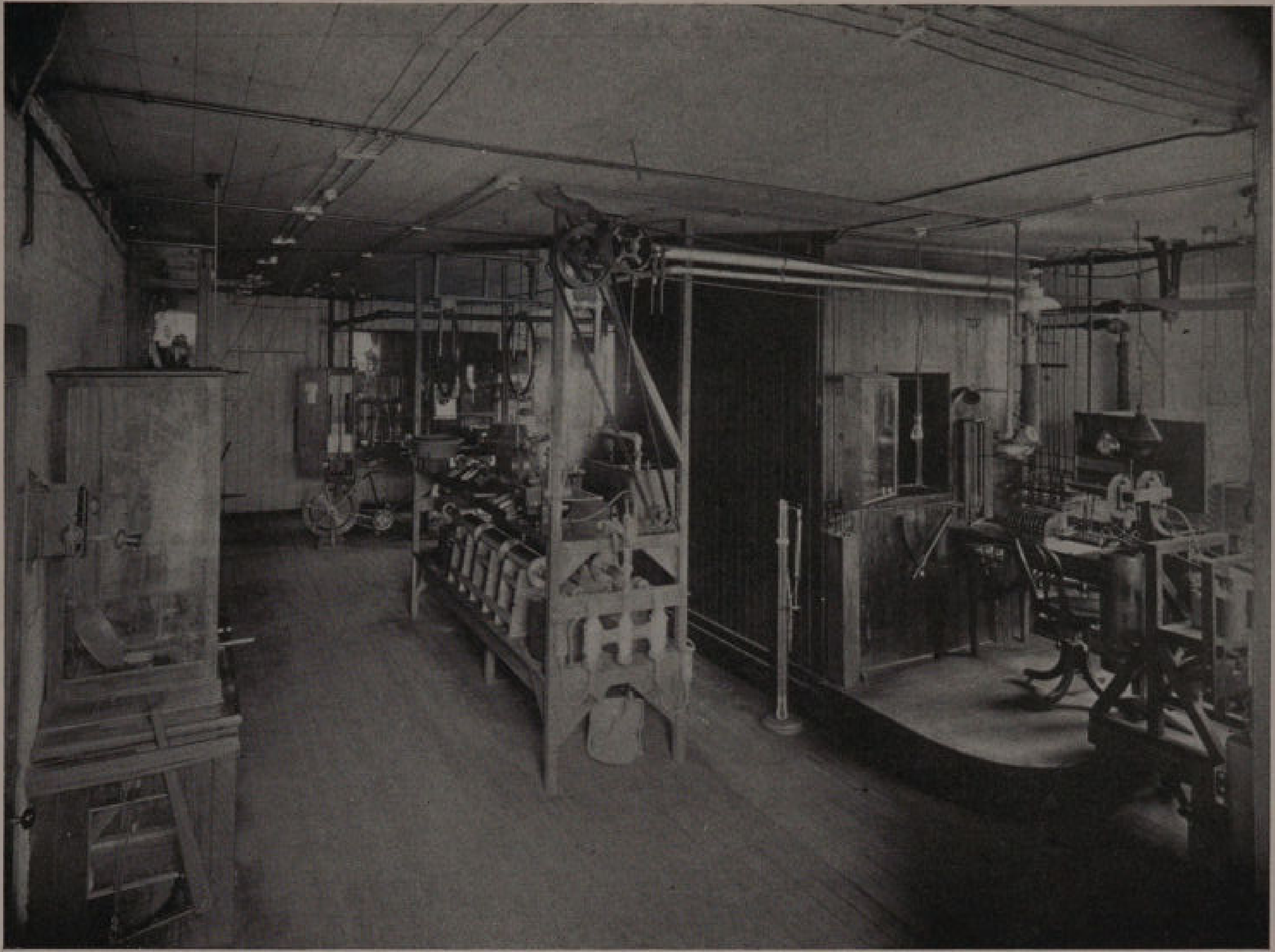


FIG. 70.—A GENERAL VIEW OF THE RESPIRATION CALORIMETER LABORATORY.

FIG. 71.—A DEVICE FOR THE REMOVAL OF THE RESPIRATORY PRODUCTS FROM THE RESPIRATION CHAMBERS. (SIDE VIEW.)

much of the protein leaves the body as comparatively complex substances, urea and the like, which in the calorimeter undergo oxidation.

Food requirements.—By the use of various forms of calorimeters, by the examination of excreta by various experiments, and by investigations of the amount of food eaten by healthy individuals in different countries and engaged in various occupations, it has been found that the amount of food that an individual requires is dependent chiefly upon the rate at which oxidation occurs in the body, and variations in this are largely due to differences in the muscular contraction that occurs. This, in turn, is dependent upon such things as the mode of life, age, sex, climate: for certain occupations and exercises entail more muscular contraction than others; cold increases muscular contraction, as is shown by the phenomenon of shivering, and it tends to induce an individual to exercise and make active movements, while warmth has the opposite effect; a strong person will expend more energy in his movements than a weak person, and a large person has a greater area of muscular surface to undergo contraction than a small person.

As oxidation uses up the fuel and muscular contraction increases the rate of oxidation, the greater the amount of muscular contraction an individual's conditions and manner of living entail, the more food will he require. It has been found, by the means mentioned in the preceding paragraph, that a man of average weight (about 59.5 kilograms—155 pounds) requires under differing conditions the following amounts of food:

When at rest, enough to yield 2000 calories.

When engaged in a sedentary occupation, enough to yield 2700 calories.

When doing a moderate amount of muscular work, enough to give 3400 calories.

When engaged in hard muscular work, enough to give 5000 to 6000 calories.

A woman of equal size, engaged in an equally arduous occupation and expending an equal amount of energy as a man, would require the same amount of food, but as the average woman is rather smaller and usually exerts somewhat less energy in her movements, it has been estimated that women require about 0.8 the amount of food that men do.

Though children require less food than adults, they need more in proportion to their size, because material is required to build their bodies as well as to repair waste and yield heat and energy. The following estimate of the food requirements is one that is often given:

A	boy	of	14	to	16	requires	about	2500	to	3000	calories
“	girl	“	14	“	16	“	“	2200	“	2500	“
“	child	“	10	“	14	“	“	1800	“	2200	“
“	“	“	6	“	10	“	“	1400	“	2000	“
“	“	“	2	“	6	“	“	1200	“	1400	“
“	“	“	1	“	2	“	“	900	“	1200	“

The comparative proportion in which the food principles are used is about as follows: sufficient protein to give about $\frac{1}{5}$ of the calories required, sufficient fat to yield a like number, and sufficient carbohydrates to make up the remaining $\frac{3}{5}$ of the calories. Thus the food for a meal intended to furnish 1000 calories would be portioned as follows: Enough protein and fat to give 200 calories each and sufficient carbohydrates to yield 600 calories. As each gram of protein and carbohydrate furnishes 4 calories and each gram of fat 9, this would mean 50 grams of protein (200÷4=50), 22.2 grams of fat, and 150 grams of carbohydrate.

volatile protein matter may be exhaled from the lungs. The kidneys are also the main channels for the elimination of foreign matter, such as drugs, toxins produced by bacteria, products of intestinal putrefaction and defective metabolism. If there is any abnormal exudate present in any part of the urinary organs as the result of inflammatory conditions some of it will be washed out with the urine, and if, as sometimes occurs when the kidneys are diseased, the kidney cells are not acting properly, substances that are normally not extracted from the blood may be present in the urine and, under such circumstances, there is likely to be a smaller amount than normal of the usual waste matter in the urine.

Thus, analysis of the urine is an aid in discovering defective conditions in the intestines, defective metabolism, and diseased conditions of the urinary organs.

Quantity, Composition, and Characteristics of Normal Urine

Quantity.—The average amount of urine voided in 24 hours by a healthy adult is about 40 to 50 ounces; by a child of

2 to 5 years, 15 to 25 ounces.
5 to 9 years, 25 to 35 ounces.
9 to 14 years, 35 to 40 ounces.

Even in health, however, the amount of urine voided may vary very considerably. The principal points to consider in deciding if difference in quantity is normal or abnormal are the amount of liquid that has been drunk and the quantity of water that has left the body by other channels, for the system tends to maintain its natural percentage of fluids and to get rid of excess. Thus, if much liquid is drunk, more will be eliminated than if only a small amount is taken, and if an unusual amount of water leaves the body as the result of vomiting, of watery evacuations from the bowels, or of excessive perspiration, there will be less to pass through the kidneys. In fever it is to be expected that the amount of urine will be diminished, partly as the result of the increased rate of evaporation of sweat, due to the high temperature. When the nervous system is depressed, as in shock, there may be either suppression or retention of urine. In some diseases of the kidneys, the secretion of urine is diminished, and in diseases such as diabetes mellitus, diabetes insipidus, and hysteria, secretion is increased; nervousness also often causes an increase in the amount of urine voided. Conditions causing a loss of fluid from the body are usually characterized by excessive thirst, which is another example of the body's endeavor to maintain its normal fluid content. The sensation of thirst is the result of the effect which the loss of water from the body protoplasm produces on some part of the nervous system.

Physical properties of urine.—Normal urine is a transparent, yellowish or light amber-colored liquid with a characteristic odor, a slightly alkaline reaction, and a specific gravity of 1.012 to 1.030, 1.020 being the average.

The specific gravity shows the relative proportion of water and solid matter constituting the urine; the greater the amount of solid matter present, the higher will be the specific gravity. This and the means of ascertaining the specific gravity of liquids were discussed in Chapter III. Even in health there can be considerable variation in the specific gravity of the urine, for, as a rule, when only a small amount of urine is secreted as the result of loss of water through other channels, as, for example, through the sweat glands, the urine will contain a relatively high percentage of solid matter and will have a high specific gravity; if, on the other hand, a large quantity of urine is excreted as the result of copious drinking or of diminished perspiration the urine will be dilute and its specific gravity will be low. In chronic nephritis, however, even though the secretion of urine is diminished its specific gravity may be as low as 1.004, since it is not only the water but also solid matter that is not secreted in the usual amounts, and in diabetes mellitus, though a large amount of urine is voided, its specific gravity is high, due to the presence of sugar.

The color of urine varies somewhat with its degree of concentration, the color being deeper when the amount of solid matter is in excess, and paler when there is a large quantity of water. Any great difference in the color or transparence of urine is usually due to the presence of foreign matter such as sugar, bile, blood, drugs, etc.

The acid reaction characteristic of human urine is due to the amount of foods that are eaten the end products of which are acid. If urine stands for any length of time it becomes alkaline as the result of the conversion of the urea into ammonium carbonate due to bacterial action. The same change may take place in the urine while it is in the urinary organs when they are diseased, and the urine is then alkaline when it is voided.

The composition of urine.—The general composition of urine is about as follows:

Water		967.	parts
Solids:			
Urea		14.230	"
Nitrogenous extractives such as	creatinin xanthin hypoxanthin	10.	"
Mucus, pigments, ferments			
Salts such as	urates sulphates phosphates chlorids	8.770	"
Gases, principally nitrogen and carbon dioxid.			

The urea, which is the principal solid constituent of the urine, is the form in which the greatest amount of waste nitrogen derived from protein metabolism is excreted from the body. The average amount excreted daily, under normal conditions, is about 40 grams. The quantity will be increased by a high protein diet and decreased by a carbohydrate diet. It is increased in diabetes and decreased in some diseases of the liver.

Creatinin.—Is thought to be formed in the body as the result of the metabolism of its muscle tissue.

The uric acid and extractives such as adenin, xanthin, and hypoxanthin vary in amount according to the quantity of nucleo-proteins taken in the food, therefore it is considered that they represent the products of the metabolism of nucleo-proteins of the body and of food. Practically all the uric acid present in urine is in the form of salts (urates).

Traces of mucus are found in normal urine, but the presence of large amounts is indicative of inflammatory conditions of the urinary organs, especially the bladder.

Abnormal Constituents of the Urine

The substances indicating abnormal conditions which are most frequently found in the urine are: Albumin, glucose, acetone, diacetic acid, indican, bile, blood, casts, calculi, pus, and bacteria.

Albuminuria is the name applied to the condition in which any form of heat-coagluable protein is present in the urine, and the protein is designated as albumin, though, as a rule, it is partly globulin. The continued presence of such protein in the urine is usually due to disease of the kidneys or to abnormal pressure changes in the renal blood-vessels, such as occurs in diseases of the heart.

Glucose may be present in the urine temporarily after the intake of a large amount of carbohydrates, especially sugars, and sometimes during pregnancy and lactation, in concussion of the brain, and during convalescence from febrile diseases, but its continued presence is always due to diabetes mellitus.

Acetone and diacetic acid are formed in the tissues chiefly as the result of the defective oxidation of the fatty acids arising during the metabolism of fats, but sometimes they develop from certain amino acids as the result of the defective metabolism of proteins. These substances are found in the urine in advanced stages of diabetes; when metabolism is retarded, as in severe anemias; and when body tissue is metabolized at an abnormally rapid rate as in starvation and in long-continued febrile diseases. Diabetic acidosis and coma are due to an accumulation of these substances within the body.

Bile is present in the urine as the result of some obstruction to its discharge from the liver or gall bladder in consequence of which it is absorbed by the blood.

The presence of **pus** is due to a suppurative inflammation of some of the urinary organs.

Blood may be present in the urine as the result of lesions in any of the urinary organs (this condition is known as *hematuria*), or disintegrated blood corpuscles may be present as the result of hemolysis resulting from any of the various causes of this condition, *e. g.*, the toxins of certain of the infectious diseases, transfusion of blood, snake bite. This condition is known as *hemoglobinuria*. In hematuria, the urine has a characteristic smoky appearance. If the blood comes from the bladder or genital organs it is likely to be clotted.

Calculi or stones are the result of the precipitation of some of the solid constituents of the urine. They may form in any part of the urinary system and are of various sizes and shapes. The precipitation is usually the result of abnormal changes in the reaction of the urine or the presence of foreign substances.

Casts consist of deposits of different kinds of substances, such as pus, blood, fatty matter, etc., that have become hardened in the kidney tubules and are washed out with the urine. They are called *casts* because they retain the cast or form of the tubules in which they were deposited. Their presence in the urine, which is ascertained with the aid of the microscope, shows some abnormal condition of the kidney from which they came.

Urine Analysis

A complete urine analysis includes taking the specific gravity of the urine, testing its reaction, determining the total quantity of solids and the percentages of the various solids, such as urea, extractives, and salts. As a rule, nurses need to know only the tests for substances, such as albumin, glucose, acetone, and diacetic acid, which, in certain diseases, it is necessary to keep track of from day to day, but a few other tests are described here since it is interesting to know how the data contained in the urine-analysis reports sent from the laboratory are obtained.

The method of testing the specific gravity of liquids was described in Chapter III. and the manner of ascertaining their reaction was mentioned in Chapter XIV.

Methods of determining the total quantity of solids.—Stir the specimen of urine, or else shake the flask containing it, so that if any sediment has precipitated it will become mixed with the liquid portion of the urine. Then pour 25 c.c., or other definite measure, of the urine into an evaporating dish, place this in a water bath over a flame, and let it remain until the water of the urine has evaporated and the residue is perfectly dry. Weigh the dish with the residue and deduct the weight of the former from the total. The amount remaining will be the weight of the residue and will show the quantity of solid matter in the amount of urine taken.

When only an approximate estimate of the total solids is required, it is often calculated by multiplying the two last figures of the specific gravity by 2.33. Thus, if the urine has a specific gravity of 1.025, there will be 58.25 grains of solid matter contained in 1000 grains of the urine; since 25 × 2.33 = 58.25.

A test for acetone.—To 10 c.c. of urine, in a test tube, add about 20 drops of concentrated NaOH and 5 drops of a 5% solution of sodium nitroprussid. If acetone is present, the mixture becomes red and the red changes to purple upon the addition of an excess (*i.e.*, a larger quantity than there is of the mixture) of glacial acetic acid.

A test for diacetic acid.—Add, drop by drop, some ferric chlorid solution to about 5 c.c. of the urine until a precipitate ceases to form; filter the mixture, to remove the precipitate, and add some more ferric chlorid to the filtrate. If diacetic acid is present a deep red color will develop. As other substances, such as those excreted after taking salicylic acid and similar drugs, give the same reaction, the solution is boiled, and if the color is due to diacetic acid, it will disappear, the acid being thereby changed to acetone, but color due to the other substances will be permanent.

Albumin tests.—The heat test for albumin was described on page 284. The Heller test, another of the tests in common use, is as follows: Pour about 3 c.c. of concentrated nitric acid into a test tube and, with a pipet, add, slowly, so as to avoid mixing the two liquids, about 3 c.c. of filtered urine. A white ring at the junction of the fluids shows the presence of albumin in the urine.

Tests for glucose.—Several of the common tests for glucose were given in Chapter XVIII.

A test for indican.—Put about 4 to 6 drops of a 1% solution of potassium permanganate into a test tube, add between 1 and 2 c.c. of chloroform, then 10 c.c.

of hydrochloric acid, and lastly 10 c.c. of the urine. Invert the test tube two or three times so as to thoroughly mix its contents and allow it to stand for about 5 minutes. If indican is present in the urine, it will be broken down by the hydrochloric acid and oxidized by the potassium permanganate to indigo, which is dissolved in the chloroform and gives a blue color, the shade of which varies with the amount of indican present, being very pale if there is only a slight trace.

A test for bile.—One of the tests for bile very commonly used is that known as *Gmelin's test*, which is as follows: To 5 c.c. of concentrated HNO_3, in a test tube, add 2 to 3 c.c. of urine, pouring the latter down the side of the tube so that the two liquids will not become mixed. If bile is present, circles of various colors will form where the liquids come in contact.

A test for blood (Heller's test).—Add enough NaOH to 10 c.c. of urine to make the latter decidedly alkaline, boil this, and then let it stand. If blood is present the alkali converts the pigment into hematin which is precipitated as a red deposit. As there are other substances sometimes present in urine that give the same reaction, filter the liquid and pour a little acetic acid over the precipitate. If this consists of substances other than hematin, it will be entirely dissolved; hematin may be slightly, but it will not be wholly, dissolved.

A test for pus.—Before testing urine for pus, let it stand undisturbed so that any sediment present will fall to the bottom, then pour off as much of the overlying liquid as possible without disturbing the sediment, and to the latter add some concentrated

NaOH or KOH solution or some strong NH_4OH. If there is pus in the sediment, it will be converted into a viscid mass. If mucus is present, it will be curded, but it will not assume the tough, gelatinous appearance that the pus does.

GLOSSARY

Alloy, a mixture of two or more metals of differing nature and value.

Anhydrid, an oxid which unites with water to form an acid; a substance formed from a compound, especially an acid, by the loss of a molecule of water.

Anion, a negative ion.

Anneal, to render a substance less brittle by heating it to a high temperature and then cooling it slowly.

Anode, the positive pole of an electric cell.

Apposition, in contact with.

Arbitrarily, depending on the will of an umpire, judge, etc.

Binary, composed of two elements.

Calcareous, pertaining to or of the nature of lime.

Cathode, the negative pole of an electric cell.

Cation, an electro-positive ion.

Caustic, corrosive.

Colloid, glutinous; resembling glue; a non-crystalline substance.

Commercial, as ordinarily furnished by dealers.

Conifers, trees that bear cones.

Decomposition, the separation of a substance into simpler parts.

Deflagration, sudden, rapid burning.

Deflected, bent from a straight line; turned aside.

Degeneration, changing from a higher to a lower form.

Dehydration, removal of water from a substance.

Deliquescence, see page 222.

Detergent, cleansing.

Diaphragm, a partition, such as a membrane stretched between two cavities.

Dissociation, partial decomposition.

Ductile, pliant; easily stretched.

Ebullition, the bubbling of a liquid caused by escaping gas.

Effervescence, the bubbling of gas in a liquid.

Efflorescence, see page 222.

Electrolysis, chemic decomposition produced by electricity.

Filament, a fine thread or thread-like process.

Filtrate, the liquid which passes through a filter.

Forge, to form into shape by heating and hammering.

Fuse, to melt.

Germination, sprouting.

Habitat, the natural locality of animals, plants, etc., in their wild state.

Heliotherapy, the treatment of disease by exposing the body to light.

Hepatitis, inflammation of the liver.

Hydrate, any compound of hydroxyl with a radical; a compound containing water of crystallization. See page 223.

Hydrated, combined with water.

Hydrolysis, chemic decomposition in which a compound breaks up as its molecules combine with one or more molecules of water.

Ignite, to set on fire.

Incandescent, glowing, white with heat.

Interaction, chemical action in which there is double decomposition and combination.

Ion, an electrically charged atom; see Chapter VII.

Isomerism, see page 158.

Nascent, beginning to grow or exist; an active state.

Neutralize, to render inactive; to make neither acid nor alkaline.

Nitrogenous, pertaining to or containing nitrogen.

Polymer, one of a series of compounds of which the molecular weights of the members of the series are in numbers which are multiples of each other.

Precipitate, to render a soluble substance insoluble and thus cause it to come out of solution and fall to the bottom of the containing vessel; the solid thus precipitated.

Radical, see page 150.

Reaction, a chemical change.

Residue, that which remains after other substances have been removed.

Rhythm, harmonical; periodical.

Sedimentation, settling to the bottom of a liquid.

Serrated, notched like the edge of a saw.

Ternary, made up of three elements or radicals.

Terpenes, hydrocarbons of the formula $C_{10}H_{16}$; they are derived chiefly from volatile oils and resins.

Volatile, tending to evaporate rapidly.

Warp, the threads which extend lengthwise in the loom and are crossed by the woof.

Welded, united by hammering or heating.

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INDEX OF EXPERIMENTS

The Tuberculosis Nurse

By Ellen N. La Motte, R.N.

308 pages. 120 Illustrations. $1.50 net.

Miss Lavinia L. Dock, author of *Materia Medica for Nurses,* etc., whose nursing and settlement work give her special authority as a judge of the value of the book, makes regarding *The Tuberculosis Nurse* the following comments:

"The book impresses myself and my associates as being an excellent—it is fair to say a masterly—piece of work. The author is a woman of unusual abilities and one could be fairly assured that a book from her would be well done. I think that her treatise should prove exceedingly valuable and useful to workers in the matters considered by her. It will undoubtedly be made of service by boards of managers who are directing crusades against tuberculosis and by nurses who are engaged in the work or who are in training for the work. Tuberculosis is important not only as a disease, but in its social aspects, and it is on this ground that the investigation of the subject is calling for such widespread attention. In its social relations it is to modern society what leprosy was to the communities of the Middle Ages. This fact will help to emphasize the importance of a book like Miss La Motte's, who has studied the subject from the point of view of preventive service."

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