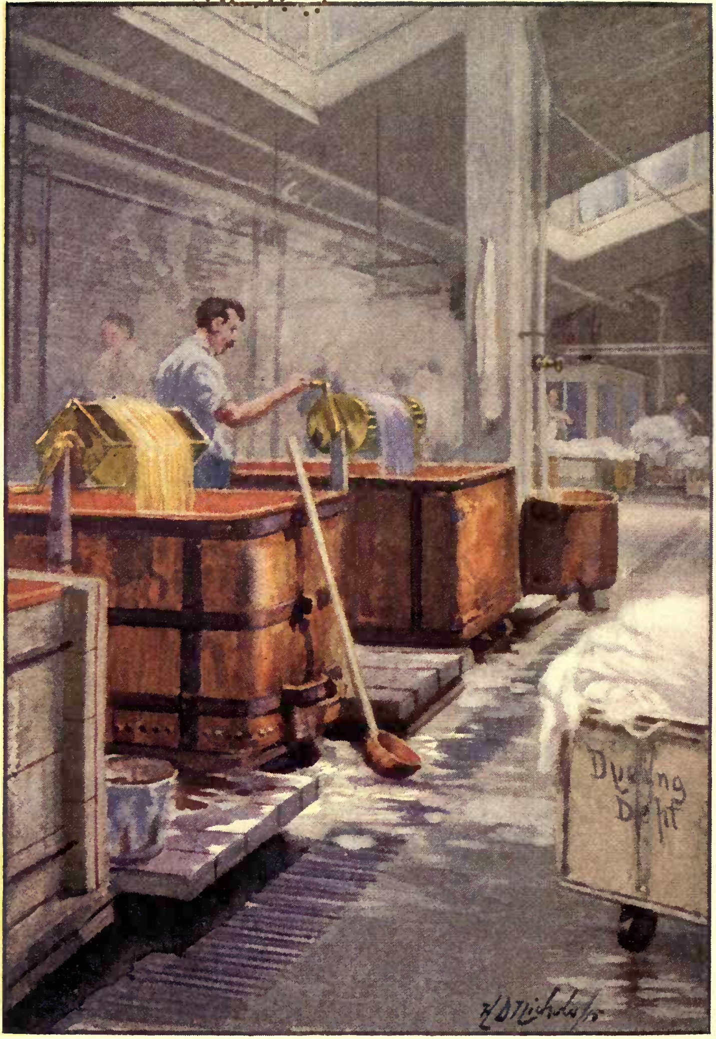
CHEMISTRY IN THE HOME

WEED



DYEING SILK.

CHEMISTRY
IN THE HOME

BY

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AMERICAN BOOK COMPANY
NEW YORK CINCINNATI CHICAGO

F. L. Griffin

WEED. CHEMISTRY IN THE HOME
W. P. I

Education

PREFACE

This book is the product of an effort to meet the needs of students who elect chemistry early in their high school course. It aims to train students in scientific thinking and to give them a fund of information concerning the chemistry of everyday things, related to industries and the home.

Much of the theory common to high school chemistries has been omitted, because it has been found to be unnecessary, in fulfilling the purpose of this text. Constant effort has been made to keep the language and style simple and to select subject matter suitable for boys and girls in the lower classes of the high school. The book represents the result of many trials and much elimination from an overcrowded field of subject matter.

It would have been impossible to prepare the book without the efficient aid of my colleagues of the Manual Training High School of Brooklyn, New York. I wish especially to thank Dr. Charles D. Larkins, the principal of the school, to whose suggestions the book is primarily due, Mr. Charles Germann, who has read the proof and made numerous suggestions, Dr. William Lamb, Mr. Mattuck, Mr. Holly, and Mr. Foster. Mr. Frank Rexford, of Erasmus Hall High School, Brooklyn, has allowed the use of his food tables for use in the food chapter, and assisted in the preparation of the chapter. Miss J. Jenness, of the Girls High School of Brooklyn, has used the notes in her classes and has made many helpful suggestions.

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CHEMISTRY IN THE HOME

CHAPTER I

PHYSICAL AND CHEMICAL CHANGES

Importance of science. Every thinking person realizes that, in order to get the most out of life, it is necessary to understand the laws that govern the world in which we live. The consequences of not understanding, or of disobeying, these natural laws cannot be evaded. The State may make a law ordering us to pay a tax upon bonds which we may own. An unscrupulous person might evade the payment of this tax and thus prevent the carrying out of the law. It is not so with the laws of nature. Nature never employs courts and officers to carry out her decrees. If we violate her laws, the penalty is certain to follow, hence the importance of understanding and obeying these laws. Moreover, the laws of nature, unlike man-made laws, never change, so that what we learn of them in our study of chemistry, will be of use to us throughout life.

We must not think of physics and chemistry as studies useful only to those who expect to engage in technical pursuits, but as everyday studies that will help us to live more satisfactory lives, because they help us to understand the reasons for the numerous things that we must or must not do if we expect to secure the greatest amount of health and enjoyment from life.

Matter is continually changing. As we consider our everyday life, we see that one of the most striking things about it is, that there is everywhere continual change. Our clothes wear out, iron rusts, trees decay and die, even the rocks that seem so everlasting crumble and change to soil.

Our world, then, is not an unalterable mass, but is in a constant state of change, and what we must do is to learn to understand and direct these changes, so that they may be for our benefit, and not for our injury. This, man is slowly learning to do. Until we had learned the cause of decay, we could not know how to prevent it, but now that we understand the reasons why foods spoil, we can prevent such changes. Scientists are constantly studying these changes which take place, and finding out how to use them to our advantage. When we examine these various changes carefully, we find that they are divided into two great classes, physical and chemical.

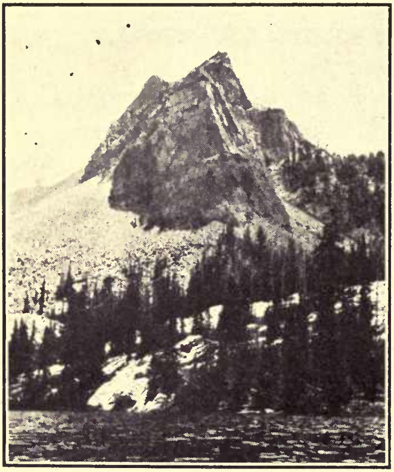

FIG. 1. — Crumbling rock makes soil.

Physical changes. When you cut out and make a suit, you have changed the shape of the cloth. You have taken cloth, thread, buttons, and lining. You have changed the relative arrangement of all these materials, but every particle of the cloth, thread, buttons, etc., is still in existence as cloth, thread, buttons, etc. When the suit is worn, you brush against the furniture, and tiny pieces of the cloth are worn off, but they are essentially the same cloth particles. Such changes as these are called *physical changes*.

Mixing a batter, boiling water, and the falling of rain are other examples of physical changes; in general the science of physics is the study of such changes. We see ourselves reflected in mirrors, our trolley cars run, we telegraph and telephone, by taking advantage of the laws governing such physical changes. Physical changes are those in which the composition of the small particles of the original substance remains unchanged.

Chemical changes. If we tear up a newspaper, we illustrate a physical change, since every particle of the original paper is still in existence. If, however, we set fire to the paper, a different kind of change takes place. The paper particles disappear, and new substances (a white ash and an invisible gas), having new properties, take their place. Such a change as this is a *chemical change*, and the study of such chemical changes and the laws which govern them, constitutes the science of chemistry.

Importance of chemical and physical changes. It is difficult to overestimate the importance of the study of these chemical changes. The beautiful colors of silks and ribbons, that so delight the eye, are produced by dyes extracted by a chemist from black, vile-smelling coal tar by chemical processes. The rails on which our trolley cars run are possible only because a chemist found how to extract iron cheaply from its ores by chemical changes. The proper selection of foods for the home may be made when we have learned of the chemical changes which each type of food undergoes in digestion, and the body requirements.

Often actions which we wish to study include both physical and chemical changes. In the making of bread, the mixing of the flour and water is a physical change, but the action of the yeast in making the bread light is a chemical change, so that the baked loaf is a result of both physical and chemical changes. Every day, in a hundred different ways we are making use of both physical and chemical changes. Let us at the very beginning try to understand and appreciate the importance of the study of these physical and chemical changes in their relations to our everyday life.

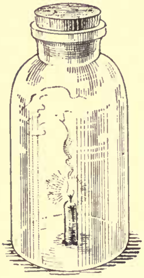

FIG. 2. — Matter is not destroyed in burning.

Matter is indestructible. In many of these physical and chemical changes, there seems to be a destruction of matter. When a candle burns, it disappears, and our natural conclusion is that the matter composing it has been destroyed. It is, however, not safe for us to jump at such a conclusion. We know that our impressions are not always to be relied upon. You know how difficult it is to determine by your feelings, whether a room is at the proper temperature or not, and how necessary it is to correct your impression by consulting a thermometer. You know how impossible it is to estimate the size of a room, or to guess a person's weight accurately; but with a yardstick you can measure the room and with a scales you can find a person's weight.

You must always try to find a way to test the truth of your impressions, and, in the case of the candle, the test is easy. If you place a candle in a large bottle filled with air, seal the bottle, weigh it, and then set fire to the candle without opening the bottle (which can be done by the aid of a burning glass), you are sure that no matter can either get into or escape from the bottle (Fig. 2). The candle will burn for a time, and then go out. If you allow the bottle to cool, and then weigh it again, you will find that its weight has not changed. That is, the materials making up the candle have formed new combinations with the air, and, these new substances, being colorless gases, usually escape our notice. No matter has been destroyed, and no matter has been created, but new substances have been formed, the total weight of which is the same as that of the original candle and air.

Similar experiments verify the fact that the total weight of the new products is always equal to the total weight of the original substances. We can neither create nor destroy matter. All we can do is to change its combinations. This is the Law of the Conservation (or Indestructibility) of Matter: *Matter can neither be created nor destroyed.*

SUMMARY

Matter is anything that occupies space.

A physical change is a change in which no particles of matter different in composition from the original substance are formed.

A chemical change is one in which the particles of matter in the original substance are so altered that the identity of the individual particles is destroyed.

Law of Conservation of Matter: Matter can neither be created nor destroyed.

Exercises

1. Name three common chemical changes. Three physical changes.

2. Are the following changes chemical or physical? Rusting of iron. Striking a match. Beating an egg. Making coffee. Sweeping. Digesting food.

3. A ton of coal when burned will produce only about 200 pounds of ash. Is this not a case of the destruction of matter? Explain.

4. Iron rust weighs more than the iron from which it was made. Has not matter been created? Explain.

CHAPTER II

WATER

Occurrence of water. Water, a tasteless and odorless liquid, is one of the few indispensable substances of the earth. It is found even in places where no visible trace of its presence exists. The paper on which these words are printed contains about 8 %, the solid rocks contain a small amount, and our foods are largely composed of it. The table below shows the amount of water present in some common substances.

Bananas	75 %	Eggs	74 %
Beef, rib roast	44 %	Our bodies	65 %
Cabbage	78 %	Potatoes	78 %
Cheese	31 %	Wheat flour	12 %

Properties of water. When we cool water to 32° F.,[1] it changes to a transparent solid, colorless in small, and light blue in large, masses. Ordinary ice is not perfectly clear, because the water from which it is made contains air dissolved in it, and, on freezing, this air is distributed through the ice in the form of small bubbles. The freezing point of water is the same as the melting point of ice, namely, 32° F.

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[1] Our ordinary house thermometers are graduated on what is known as the Fahrenheit scale. Fahrenheit is usually abbreviated to F. The name is given in honor of Dr. Fahrenheit, a German, who devised the scale.

Expansion of water. Most of us are familiar with the fact that solids expand when heated. The next time you cross an iron bridge, see if you can find the expansion joint provided to allow for the expansion of the iron. Such expansion joints are used in many places to prevent the straining of the metal that occurs when it is held rigidly and heated. There are many such expansion joints in the steam lines of your school building. Figure 3 shows a section of one of these expansion joints used in long pipe lines for steam and hot water. See whether you can find one.

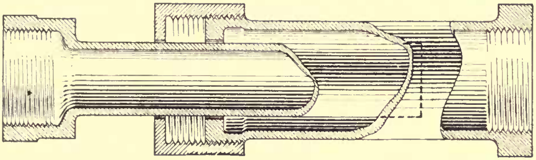

Fig. 3. — An expansion joint.

The blacksmith takes advantage of this expansion of iron in putting the tire on a wagon wheel. He makes the tire too small to go on the wheel while cold, but, on heating the tire, it expands sufficiently to allow it to be forced on. When the tire cools again, it shrinks and grips the wheel so tightly that it stays on without bolts or screws to hold it.

Water expands in the same way when heated, but there is an unusual fact about its expansion. If we start with water at 32° F., its freezing point, it *contracts* on heating, up to the temperature of 39° F., at which point it first begins to expand. At 39° F., then, water has its least volume, and therefore its greatest, or maximum, density.[1] That is, a cubic inch of

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[1] By density we mean the quantity of matter contained in a given volume of a substance. A cubic foot of water weighs 62½ pounds. A cubic foot of iron weighs 437½ pounds. These numbers express the density of water and iron respectively. Volume for volume, the

water at 39° F. weighs more than a cubic inch of water at any other temperature.

As a result, when the surface water cools, it becomes denser and sinks to the bottom. This continues until all the water has reached its temperature of maximum density, 39° F. Then, as the surface water becomes slowly cooler, it becomes less dense, and therefore floats on the surface of the warmer water beneath. The ice forms on the surface and acts as a protective blanket. Because the ice is a poor conductor of heat, the water below freezes very slowly. It is for this reason that a long cold spell is necessary before we have skating on a pond or lake. The water at the bottom of the ocean has about the temperature of maximum density, 39° F., the year round.

FIG. 4. — Expansion from freezing.

As water freezes, it expands, and it is this expansion that cracks water pipes. You have noticed the forcing up of the cap of the milk bottle in winter, due to the expansion of the milk in freezing (Fig. 4). This expansion is the reason why ice floats on the water. Can you explain why thick cut glass dishes break, if put into hot water, while thin glass tumblers do not?

Steam. If we melt ice, and heat the resulting water to 212° F., and continue the heating, the water boils, that is, changes to an invisible gas called steam.

The mist that is commonly called steam is not steam at all. Notice the spout of the tea-kettle when the water is boiling vigorously. For the space of a half inch from the end of the

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iron weighs seven times as much as the water, or a cubic foot of iron contains seven times as much matter as a cubic foot of water. The density of iron, compared to water, is seven.

spout, nothing can be seen; beyond that point we have the mist formed by the steam cooling and condensing into very tiny drops of water. That is, we have transparent steam issuing from the spout, and beyond this a cloud of condensed steam.

Water boils at 212° F. only when the pressure of the air is normal, that is, when the barometer[1] stands at 30 inches. At Denver, the ordinary air pressure is much less than at New York, owing to its elevation above sea level. Water boils in Denver at about 202° F. This makes it necessary to boil foods longer in Denver than in New York, as the temperature of the boiling water is so much lower there. On the tops of some high mountains, it is almost impossible to hard boil an egg, the temperature of the boiling water is so low.

Increasing the pressure raises the boiling point. In a steam boiler, if the pressure is 100 pounds per square inch, the boiling point of the water is raised to 337° F. Advantage is taken of this fact, in boiling meat and vegetables at high altitudes. The food, with water, is placed in a steam-tight vessel, called a pressure cooker

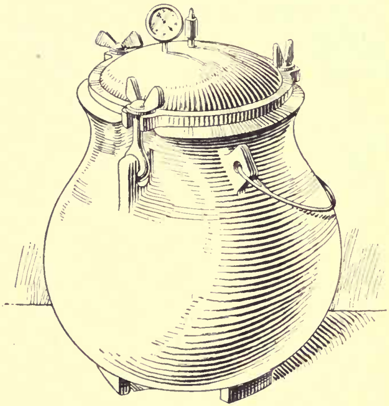

Fig. 5.—A pressure cooker.

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[1] A barometer is an instrument that tells what the pressure of the air is. This pressure varies from day to day. When it is sufficient to make the barometer stand at 30 inches, we say that the air pressure is normal. (See page 97.)

(Fig. 5), and heated. The steam, unable to escape, increases the pressure, this raises the boiling point, and so it is possible to cook the food satisfactorily.

Evaporation of water. As found in nature, water usually contains impurities. The purest form is rain water, collected after the rain has fallen for a time sufficient to wash the dust from the air. If we consider the cycle of water in nature, we shall understand better why our spring and river waters must contain some impurities.

We know that if a pan of water is allowed to stand exposed to the action of the air and the sun, the water disappears. Where has it gone? Since it does not disappear when in a closed vessel, it is inferred that it must have passed into the air. It is not visible, but exists in the form of an invisible vapor. We say that the water has *evaporated*, and we call the change *evaporation*.

This same evaporation occurs on a large scale at the surface of the ocean. Sometimes the phrase, " the sun draws the water," is used. This is not exactly true; we should say, " the water evaporates." The resulting water vapor mixes with the air. A given quantity of air cannot hold more than a certain amount of water vapor, the exact amount depending upon the temperature. The higher the temperature of the air, the more water vapor it can hold.

The air over the ocean absorbs the water vapor which rises from it. The air, laden with water vapor, may be blown inland. When the air cools, it cannot hold as much water and so some of it condenses and forms clouds. Fog and mist are due to the same cause, and are really low clouds. This condensation continues until the drops of water become so large that they fall as rain. Before it can rain, then, it is necessary that the air contain water vapor, and that this air be cooled. This explains why certain regions have more or less rainfall than others.

The water cycle. The rain, in falling through the air, dissolves certain gases from it, and washes out the dust. Notice how dirty the first rain water that falls is, especially in the city. The amount of water that falls each year in the form of rain or snow is very great. It is enough, if it were equally distributed, and fell all at one time, to make a layer about three feet deep over all the land. This amounts to about 30,000 cubic miles of water.

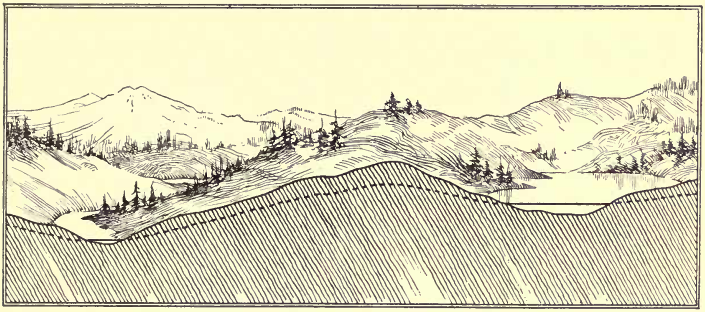


FIG. 6. — Water table. The ground is saturated with water below the dotted line.

When the rain strikes the ground, a portion of it sinks in, and slowly passes through the soil, until it meets an impervious layer, such as a bed of clay. Here it may accumulate and saturate the soil for some distance above the clay. This reservoir of water contained in the ground is called *ground water*. It is this ground water that supplies water to wells. The upper surface of the saturated soil is called the *water table*. This is not necessarily level, but may have a gentle slope. In case this water table rises as high as the surface of the ground, a spring, pond, lake, or stream is formed (Fig. 6).

The height of the water table varies with the amount of rainfall, and hence, during very dry seasons, it may become so low that springs, and even streams, may dry up and cease to run. The spring water flows into a brook, the brook into a river, and the river into the ocean. Thus the water cycle is completed. The water came from, and returns to, the ocean, and this continues over and over again.

How water becomes impure. In passing through the ground, the water dissolves any soluble matter that may be present, and thus becomes somewhat impure. Most of these impurities are harmless, but if the water comes in contact with sewage, it may be so contaminated as to be dangerous

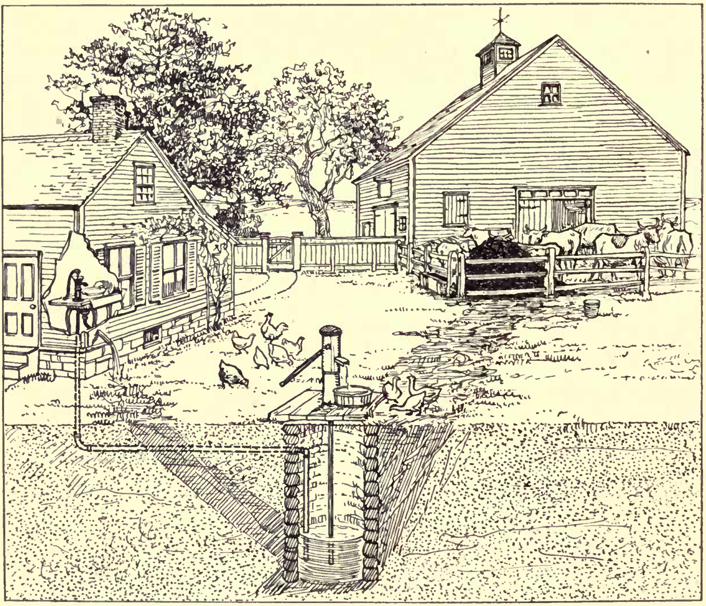


FIG. 7. — Contamination of well water.

to drink (Fig. 7). It can usually be made safe by boiling, and this is a wise precaution to take, in case there is doubt as to the purity of the water.

Distilling water. The commercial purification of water is usually accomplished by distilling it. The water to be purified is placed in a vessel (still), closed at the top, and provided with only one escape pipe. The water is heated to the boiling point. Most of the impurities, such as salt and calcium sulphate, are not volatile, and so remain in the still. Many of you have noticed the crust that is formed on the inside of the tea-kettle. This is the result of the evaporation of the water, leaving the dissolved matters behind. The water in the still (*A*, Fig. 8) changes into steam, and this steam passes through the escape pipe, which is surrounded with cool water in the condenser *C*. Cooling the steam condenses it, and pure water drops from the end of the pipe into *D*. This distilled water will not be perfectly pure, as it still contains some dissolved gases, but it is so nearly pure, that it is used in the laboratory for chemical work.

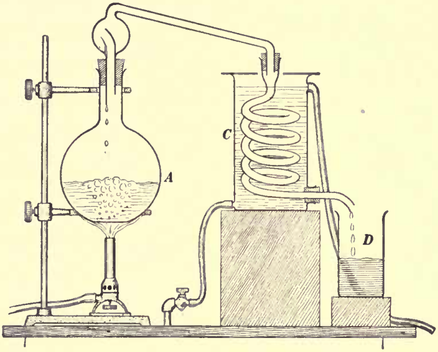


Fig. 8. — Distillation of water.

This same process, distillation, is used to separate two liquids having different boiling points. Alcohol, when it is made, is a very dilute solution of alcohol in water. This mixture is distilled, and as alcohol has a lower boiling point than water, 173° F., it boils first, and on condensing the vapor, alcohol containing only a small percentage of water is obtained. This process, called fractional distillation, is also used to separate crude petroleum, as it is pumped from the earth, into gasoline, naphtha, kerosene, and lubricating oils.

Purifying water. Water as it comes from the ground often contains suspended matter. This is especially true of the waters of the western rivers. This suspended matter may be removed by filtration, or by allowing the water to stand until the suspended matter has settled. The water supply of cities must sometimes be taken from muddy rivers. It may be cleared by filtering it through large beds of sand. This removes the suspended matter, but does not remove the bacteria, nor the dissolved matter.

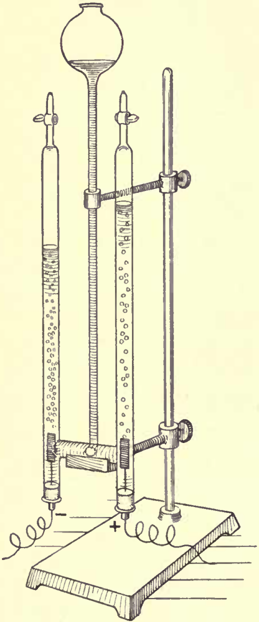


Fig. 9. — Electrolysis apparatus.

Composition of water. Water was long thought to be an *element*, that is, a substance so simple in its composition that it could not be broken up into other substances. We now know that this is not true, but that water is composed of two gases combined chemically.

We determine its exact composition by means of *electrolysis*. Pure water will not conduct the electric current, but if a little sulphuric acid is added, it allows the current to pass. This acidulated water is placed in a U-shaped tube, so arranged that an electric current can be sent through the water. The electricity breaks up the water particles, and two gases are set free, one in each arm of the U tube (Fig. 9). There is twice as much of one gas as of the other.

The gas present in the larger volume is named hydrogen, and is noteworthy because it is the lightest gas known, and because it burns with an exceedingly hot flame. The other gas, oxygen, is one of the gases contained in the air, and it is essential to life.

When we pull any substance apart in order to find out its composition, we call the operation *analysis*. By passing the electric current through the water, we analyze it, and find its composition by volume to be two volumes of hydrogen to one volume of oxygen.

Elements. Chemists have analyzed many thousands of different substances to find of what they are composed. The results are somewhat surprising. They show us that all the materials that form our world are made up of only about 83 different kinds of matter, and that by combining these in different proportions, we may make all of the various substances that we know. Thus, sugar is found to contain three different kinds of matter, that is, three elements. They are oxygen and hydrogen, the elements that we found in water, and carbon, which is the name chemists give to the black material in coal.

Occasionally a new element is found. Radium is an example of this. It was discovered in 1898 by Madame Curie.

Notice that the definition of element is not a substance that *cannot be*, but a substance that *has not been* decomposed into simpler substances. In the past, materials have been thought to be elements, that later have been decomposed. Of the 83 elements known we shall study only a few of the more common ones, as oxygen, hydrogen, chlorine, iron, and sodium.

Compounds and mixtures. When elements combine to form pure, definite substances, we call them *compounds.* Sugar, salt, and starch are examples. If elements or compounds are mixed together without any chemical change taking place, we call the result a *mixture.* Sea sand is a mixture. It contains sand, salt, seaweed, etc., all mixed together, but not chemically combined. Muddy water is another example. The mud is mechanically mixed with the water, and can be separated from it by filtering.

Law of Definite Proportions. One distinction between a compound and a mixture is that the composition of a compound is always the same, while the composition of a mixture may vary between wide limits. Water has been analyzed thousands of times with the invariable result that its composition is always found to be, hydrogen $\frac{2}{3}$, and oxygen $\frac{1}{3}$, by volume. Water, then, must be a compound, for we cannot imagine that any accidental mixture of hydrogen and oxygen would always have the same proportions.

Brass is made by melting together zinc and copper. On analyzing different specimens, we find that the percentage of zinc varies from 40% to 75%. Brass, then, must be a mixture and not a compound. These facts give us the Law of Definite Proportions, which may be stated thus: *The composition of every pure chemical compound is always the same.*

Synthesis of water. We may show the composition of water in still another way. If we mix two volumes of hydrogen with one volume of oxygen, put the mixture in a glass tube, one end of which is closed, set the open end in a dish of mercury, so as to prevent any air from getting in, and then set fire to the mixture, the gases will disappear, and a few drops of water, formed by the combination of the two gases, will appear in the tube (Fig. 10). We can easily set fire to the gases by passing an electric spark between the points of platinum wires which are sealed in the glass tube, so that their ends do not quite touch.

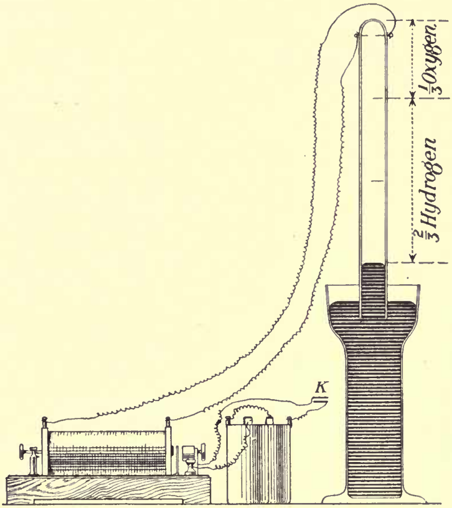


FIG. 10. — Synthesis of water

The formation of water in this way is called *synthesis*. Chemists can build up in such ways as this a great many substances that nature makes in plants and animals. Thus we have synthetic indigo, vanilla, camphor, etc. These are not imitations, but are the real substances, formed by man, instead of by nature, from the materials of which they are composed. Often such synthetic products are cheaper than the natural, and equally good. It is, of course, a fraud when they are sold as natural products, instead of being marked synthetic.

Composition of water by weight. The composition of water by weight can easily be determined. Oxygen weighs, volume for volume, sixteen times as much as hydrogen. Since water is composed of two volumes of hydrogen to one of oxygen, its composition by weight must be two of hydrogen to sixteen of oxygen, or one of hydrogen to eight of oxygen. That is, by weight water is $\frac{1}{9}$ hydrogen and $\frac{8}{9}$ oxygen.

SUMMARY

Density is the mass per unit volume.

Water *boils* at 212° F. and *freezes* at 32°F.

The composition of water *by volume* is one of oxygen to two of hydrogen; *by weight*, eight of oxygen to one of hydrogen.

An element is a substance that has not been decomposed into simpler substances.

A compound is a pure substance composed of elements chemically combined.

A mechanical mixture is a substance composed of two or more ingredients the individual particles of which are not chemically combined, but exist side by side.

Analysis is the pulling apart of a substance to find out of what it is composed.

Synthesis is the formation of a compound from the elements composing it.

Electrolysis is analysis brought about by electricity.

Law of Definite Proportions: The composition of every pure chemical compound is always the same.

Exercises

1. Is distilled water free from all impurities? Explain.

2. Is brass an element? Explain.

3. Name two common elements.

4. Is a lamb chop a compound, an element, or a mixture? Explain.

5. How can you prove that wood is not an element?

6. Does the boiling of water remove the impurities? If not, why does it make the water fit to drink?

7. How would you distinguish between distilled water and perfectly clear colorless spring water?

8. Why may water drawn from a well near a farmhouse be clear, and still unfit for drinking purposes?

9. Is synthetic indigo as good as natural indigo? Explain.

10. Why do water pipes sometimes burst in cold weather?

CHAPTER III

SOLUTION

Solution explained. We are all familiar with the change that takes place when we stir sugar in our coffee. The solid sugar disappears, the tiny particles that compose it being distributed uniformly throughout the coffee. The resulting liquid we call a *solution* of sugar.

We are so accustomed to seeing and thinking of matter in large masses, that we do not realize how small the individual particles are that make up these masses. A bit of dye, aniline violet, as large as the head of a pin, will distinctly color five gallons of water, when dissolved in it. That is, the individual particles that make up this substance are so small, that, when they are separated from each other, and mixed with the water, there are enough of them present in each drop of the five gallons to give it a distinct color.

Many substances when mixed with water behave in this way. Their particles are separated from each other and spread uniformly through the water, and, even on long standing, these particles do not separate from the water, but remain uniformly mixed with it. This intimate, uniform, and permanent mixture of a solid and a liquid we call a *solution.*

Emulsions. When oil and water are shaken together, they mix, but on standing the two separate. If, however, some mucilaginous material is added to the water, it seems to coat over the small globules of the oil, and even on long standing these do not separate from the water. Cream and milk are examples of such mixtures. We call them emulsions.

Suspension explained. Other substances when mixed with water behave in an entirely different manner. At the seashore you have doubtless watched the waves roll in on the beach. The sand mixes with the water, but does not dissolve and disappear. Instead, it soon settles, and the sea water is once more clear. We cannot produce a lasting and uniform mixture of sand and water.

Muddy water is an example of the same thing. Fine soil from the ground mixes with the rain water, and makes it turbid. It will not, however, remain uniformly mixed, but, on standing, the particles of soil slowly settle, and the water becomes once more clear. The time required for the particles of soil to settle depends upon the size of these particles. If they are very small, it may take days, but eventually the soil and water will separate. We call such a mixture of a liquid with the relatively coarse particles of a substance, a *suspension.*

The same liquid is often both a solution and a suspension. Coffee, when properly made, is a solution, but we sometimes find a sediment in the bottom of the cup, showing that the coffee was not only a solution, but a suspension as well, and that, on standing, some of the fine particles have settled.

Solution and suspension defined. A solid is in solution in a liquid when the particles of the solid are uniformly and permanently scattered through the liquid, while a solid is in suspension in a liquid when its particles in a finely divided state are mixed with a liquid, but the mixture will not remain a uniform one. Before reaching the end of the chapter, however, we shall see that these definitions need to be somewhat modified.

Filtration. To distinguish between a solution and a suspension, the easiest way is to strain, or, as chemists say, *filter* it. In the laboratory we use filter paper for this purpose. This is a porous paper that will allow water and dissolved particles to pass through, but will not permit the small particles of suspended matter to pass. If we pour muddy water through such a paper, the sediment will be held back, while the clear water passes through. The clear liquid that runs through is the *filtrate*, while the solid matter left on the paper is the *residue.* In the home, a number of thicknesses of finely woven cloth will serve the same purpose.

Many of the water filters that screw on to water faucets contain charcoal, and the water passing through its fine pores is filtered. Do not forget, however, in using such filters, that the dirt removed from the water remains in the filter, and that the filter should be frequently cleaned.

In some water filters made of earthenware, the pores of the filter are so fine that even bacteria are kept back. The finer the pores, the more thoroughly the filter acts, but also the more slowly the liquid passes through, so we must use discretion in selecting filters. The larger the pores, the better, so long as they are small enough to retain particles of the size we wish to filter out. To remove cranberry seeds and skins from the cooked mass, a fine cheesecloth is sufficient, while, to filter muddy water, a very much finer filter is necessary.

Naming different kinds of solutions. There is a limit to the amount of a substance that will dissolve in a liquid. If we add a little granulated sugar to a test tube full of water, and shake, the sugar disappears. It has *dissolved,* and we obtain a clear *solution.* If we then add a second small portion of sugar, and shake, it also dissolves. If, however, we continue to add sugar, shaking after each addition, we at last come to a point where it will no longer dissolve, but remains in the bottom of the test tube. We call such a solution, a *saturated solution* at that temperature. The water has then dissolved all of the sugar that it can. If, however, we warm the test tube, the excess of sugar in the bottom will dissolve. As we again add more sugar, we find that it also will dissolve, that is, the solution is no longer saturated. We must again add a considerable amount of sugar, to produce the condition in which some remains in the bottom of the test tube, that is, to saturate the solution at the higher temperature. On cooling, the solution will deposit sugar crystals, similar to rock candy, until the cooled solution contains the same amount of sugar that it did before heating.

It is generally true that hot water will dissolve more of a solid than cold water. It is not enough, in speaking of a saturated solution, to say simply that it is saturated; we must also state the temperature. A saturated solution of alum contains only 5 parts of alum in 100 parts of water at 32° F., while at 190° F. it contains over 200 parts in 100.

When a solution is nearly but not quite saturated, we call it a *concentrated* solution. If it contains somewhat less of the dissolved substance, it is called *strong,* while, if it contains only a little, we speak of it as *weak* or *dilute.*

A *tincture* is a solution in alcohol, such as tincture of iodine. Vanilla and lemon extracts are tinctures.

It is awkward to have to speak continually of the "dissolved substance." We therefore give this a name; it is called the *solute,* while the liquid that dissolves it is called the *solvent.* We speak of their mixture as a *solution.* Thus, our sirup of sugar is a *solution,* the sugar is the *solute,* and the water is the *solvent.*

When nothing is said as to the nature of the solvent, it is understood that water is meant. Other solvents such as ether, alcohol, gasoline, etc., are often used when the substance is not soluble in water. For instance, grease, which is not soluble in water, dissolves in gasoline.

Solutions of liquids. Not only solids, but many liquids, are soluble in liquids. When we buy alcohol it is often labeled 95%. This means that 95% of the liquid is alcohol, and 5% water. The water, a liquid, has dissolved in alcohol, another liquid. In this particular case, alcohol and water mix in all proportions, that is, we cannot have a saturated solution of alcohol in water. When this is the case, we call the liquids *miscible.* Usually, however, with liquids, as with solids, only a certain amount of one liquid will dissolve in another. We may therefore have a saturated solution of one liquid in another, as ether in water.

Solutions of gases. That liquids can dissolve gases, is shown by what happens when a glass of cold water stands for some time in a warm room. You know that the inside of the glass becomes covered over with tiny bubbles of some gas. Knowing as you do that fish require air to live, it is easy to guess that this gas is air that has been dissolved in the water. Evidently, gases do not behave as solids do, for warming the water has *decreased* its ability to dissolve air, and so the air forms bubbles on the glass.

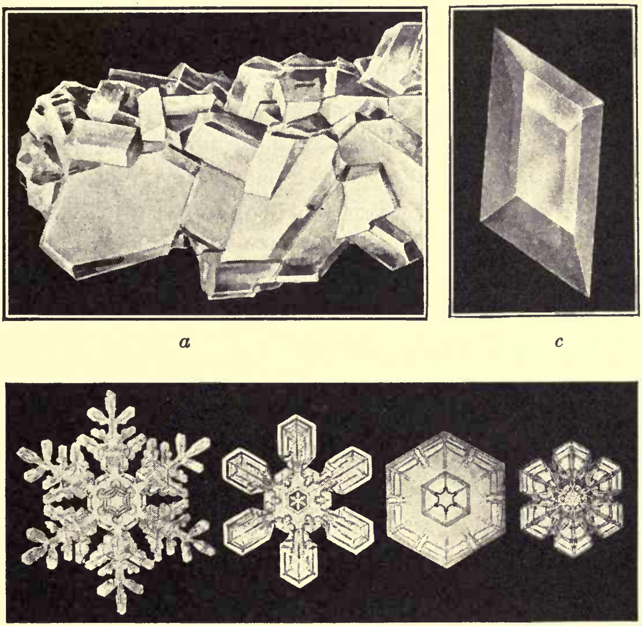
The ammonia water that you use for cleaning purposes is a solution of ammonia gas in water. Why does the cork of the ammonia bottle sometimes fly out on hot days? Soda water is another common example of a gas dissolved in water. Remember that, while heating generally increases the solubility of solids in liquids, it decreases the solubility of gases in liquids.

Effect of pressure on solubility of gases. The next time you draw a glass of vichy or seltzer from a siphon, notice that the liquid inside the siphon fills with gas bubbles, and that on standing these disappear. This is because the solubility of a gas in a liquid is dependent not only upon the temperature, but upon the pressure. When we draw from the siphon, we diminish the pressure inside. This decreases the solubility of the gas, and gas bubbles form. This goes on until enough gas has escaped so that the water is saturated at the new pressure, when the water once more becomes clear. Roughly, doubling the pressure doubles the solubility of a gas in a liquid.

Many gases are very soluble in water. Ammonia gas is so soluble in water, that more than 700 quarts of it will dissolve in one quart of water at ordinary room temperature.

Solution not limited to solids. You know that air will dissolve water, because if a pan of water is exposed to the air, the water disappears. Water will also dissolve air, for it is this dissolved air that fish breathe. You are also familiar with cases of solids and liquids dissolving in liquids. Gases also dissolve in solids, as we shall find later in our work. We must then not limit our use of the word solution to the case of solids dissolving in liquids, but we must think of solids, liquids, and gases as being all soluble in other solids, liquids, or gases. This of course does not mean that every gas, liquid, or solid is soluble in every other gas, liquid, or solid, but certain ones are soluble in certain others.

Crystallization. When a saturated solution is allowed to stand, exposed to the air, some of the solvent evaporates, and the solute is deposited in the dish in the form of *crystals*, that is, in the form of some geometrical solid, bounded by plane surfaces.


a *c*

b

Fig. 11. — Crystals. *a*, rock candy; *b*, snow; *c*, washing soda.

At the first opportunity, examine some rock candy. You will find that the pieces of sugar have a definite shape, and that each resembles its neighbor (Fig. 11, *a*).

You may have noticed the regularity of form of snow and frost crystals (Fig. 11, *b*). If you have not, at the next snowfall, catch a few flakes of snow on a black cloth and examine them under a reading glass. The crystals may then be distinctly seen.

There are many substances, such as ice, that are *crystalline,* that is, composed of crystals. These may be so crowded together that distinct separate crystals cannot be seen. When we separate crystals that have formed in a liquid from the liquid and dry them, a substance comparatively free from impurities, is obtained. Crystallization is often used to purify compounds.

Examine some salt at home with a small magnifying glass. You will find that you can make out the shape of its crystals without difficulty. The crystals of different substances have different shapes, and it is often possible to identify a substance by the shape of its crystals. Not all substances crystallize. Flour does not. We call such a substance *amorphous.*

Water of crystallization. Occasionally, crystals in forming from solution combine with water. Common washing soda is an example of this (Fig. 11, *c*). If you place a dry crystal of washing soda in a test tube and heat it, a large amount of water is given off, and a white powder remains in the bottom of the test tube. In this particular case, 106 pounds of washing soda will combine with 180 pounds of water, to form 286 pounds of crystalline washing soda. That is, more than half of the crystal is water. This water which is not mechanically mixed with, but chemically combined with, the washing soda, is called *water of crystallization.*

Efflorescent, deliquescent, and hygroscopic substances. If crystals of washing soda are allowed to stand in the air, they give up this water of crystallization, and fall to a white powder. This is called *efflorescence.* Only a few of the compounds containing water of crystallization are efflorescent.

Soda lye, or, as chemists name it, sodium hydroxide, behaves in quite a different manner. If a lump of this compound is exposed to the air, it absorbs water from the air, and becomes a solution. This is called *deliquescence.*

Many substances that are not deliquescent will absorb a little water. If paper is dried, weighed, exposed to the air, and weighed again, it will be found to have absorbed about 10% of water. We cannot call paper deliquescent, because it will not go on absorbing water until it dissolves. Instead, we call it *hygroscopic.* Cloth, wood, and leather are all hygroscopic.

Supersaturated solutions. If vinegar, or, as chemists name it, acetic acid, acts on washing soda, a chemical change takes place and a new substance, sodium acetate, is formed. This substance is readily soluble in water, but its solution shows certain peculiarities.

If we prepare a cold saturated solution of sodium acetate, add a large excess of the solid, and heat, the solid all dissolves. On cooling, you would naturally expect that the excess of sodium acetate would crystallize, and that we would again obtain a cold saturated solution mixed with crystals. Most substances do act in this way, but sodium acetate solution on cooling behaves differently, since no solid separates. By preparing a hot concentrated solution and then cooling, we have obtained a solution that contains many times as much sodium acetate dissolved as we could have dissolved by merely shaking the solid with cold water. Such a solution is called *supersaturated.*

A supersaturated solution will remain in a fluid condition, if put in a flask and corked, until a crystal of the solute is dropped into it. Then the excess of the solute over that required to form an ordinary saturated solution crystallizes, and an ordinary saturated solution, containing crystals of the solute, results. During the process of crystallization much heat is given out, and the temperature of the solution rises. A number of substances form supersaturated solutions, among others the "hypo" of the photographer.

Self-heating hot water bottle. Advantage has been taken of supersaturated solutions to prepare a hot water bag that can be used at any time without requiring hot water. A rubber bag is filled with a supersaturated solution, which, so long as it remains corked, will remain fluid. When we wish to use the bag, we take out the cork, blow upon it to evaporate the water, thus obtaining a thin film of crystals upon the cork, and then replace the cork. Crystallization then takes place and the mass becomes warm. The bag can then be used as any other hot water bag would be. Before using again, the rubber bag, corked, is placed in hot water until all of the crystals have dissolved. It is necessary to be sure that every crystal, no matter how small, has dissolved, as otherwise the supersaturated solution would again crystallize as the fluid cools. Such bags are not in common use because hot water is so easily available.

You can easily prepare such a device for home use, using a bottle containing a supersaturated solution of "hypo," or, as it is correctly named, sodium thiosulphate. To prepare such a solution, dissolve a pound of hypo in a half ounce of boiling water. If the solution is not perfectly clear, filter it through absorbent cotton, pour the liquid into a pint jar, and seal tightly.

SUMMARY

A solution is a uniform mixture of substances which do not separate even on long standing.

A suspension is a mixture of substances that separate on standing.

Defining the parts of a solution. The *solute* is the substance dissolved; the *solvent* the material in which the solute dissolves; and a *solution* is the result.

Strength of a solution. A saturated solution is one containing as much of the solute as will dissolve at the given temperature. A concentrated solution is one not quite saturated. A dilute solution is a weak solution.

A tincture is a solution in alcohol.

The filtrate is the clear liquid that passes through the filter paper.

Miscible liquids. When one liquid will dissolve in another in any proportion, the liquids are *miscible.*

A crystal is a natural geometrical solid bounded by plane surfaces.

Water of crystallization is water chemically combined in a crystal.

An efflorescent substance gives up its water of crystallization to the air, without being heated.

A deliquescent substance absorbs water from the air, finally dissolving in it.

A hygroscopic substance absorbs a limited amount of water from the air but does not form a solution.

Heat and solubility. Heating generally increases the solubility of solids and decreases the solubility of gases in liquids.

Exercises

1. Is a cup of tea a solution or a suspension? Explain.

2. Why does dipping a greasy waist in gasoline clean it?

3. Why does alcohol clean eyeglasses better than water?

4. How would you test silk, to find out if it is hygroscopic?

5. In buying washing soda, is it well to insist on having clear glassy crystals? Explain.

6. Why is soda lye put up in soldered tin cans?

7. Name three natural crystals with which you are familiar.

8. Name three amorphous substances.

9. Is ice crystalline? Explain.

CHAPTER IV

OXYGEN AND HYDROGEN PEROXIDE

Occurrence of oxygen. From the study of water you know something of the properties of oxygen, and of the great importance of this element to life. Oxygen is found everywhere. Water is by weight $\frac{8}{9}$ oxygen; our bodies are $\frac{2}{3}$ oxygen; sand is $\frac{1}{2}$ oxygen; while marble, washing soda, vinegar, meats, and vegetables contain it in considerable proportion.

That compounds containing oxygen should be so widely distributed is natural. You will remember the action of the air on various substances. The rusting of metals, the decay of wood, and the spoiling of foods, all result in the formation of oxygen compounds. These changes have always been going on, and, as a result, large amounts of such oxygen compounds now exist everywhere. It is estimated that 50% of the earth's crust is oxygen.

Methods of preparing oxygen. You have already learned that the decomposition of water by electricity yields oxygen. But when we wish to obtain it in considerable amounts in the laboratory, we resort to another method, the heating of potassium chlorate. This white crystalline compound, that you may have used in the form of tablets for the throat, contains 39% of oxygen, all of which it gives off on being heated. It has been found that the addition of another substance, manganese dioxide, causes the potassium chlorate to decompose at a lower temperature. We therefore use a mixture of these two substances in preparing the gas.

A mixture of three parts of potassium chlorate, and one part of manganese dioxide, is placed in a test tube (Fig 12) provided with a cork and a delivery tube, and the end of the delivery tube placed under water in a pneumatic trough. When the mixture in the test tube is heated, it decomposes. A gas passes off through the delivery tube and, bubbling through the water, escapes into the air. The gas which first escapes is the air that was in the test tube. As soon as all the air has been driven off, the gas which then escapes is oxygen. If a bottle is filled with water and inverted over the end of the delivery tube, the oxygen will displace the water, and will fill the bottle with pure oxygen. Many gases can be collected in this manner over water. This method is called collection by *displacement of water*.

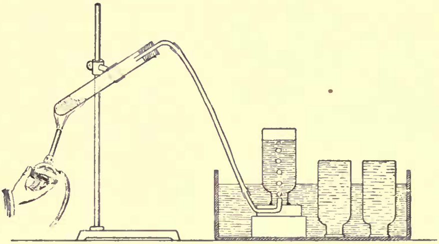


Fig. 12. — Preparation of oxygen.

Catalytic agents. After all the oxygen has been driven off, water, added to the contents of the flask, dissolves one of the residues. If the mixture is filtered and the filtrate is evaporated, a white solid different from the original potassium chlor*ate* results. It is called potassium chlor*ide*. Notice carefully the difference in the endings of the names of these two compounds.

The black residue on the filter paper is the manganese dioxide originally used. It has undergone no change. There are many such cases in chemistry where we add a substance for the purpose of aiding or checking some chemical action, in which the added substance does not itself undergo any permanent change. Such an action is called a *catalytic action*, and the substance added, a *catalytic agent.*

Properties of oxygen. An examination of the oxygen in the bottle shows that it is a colorless, odorless, and tasteless gas. We should expect this, as the air which is one fifth oxygen is colorless, odorless, and tasteless. By subjecting it to great pressure and intense cold, it can be changed into a pale blue liquid. All gases can be liquefied by this method. Oxygen is slightly heavier than air and is somewhat soluble in water, 4.1 volumes of oxygen dissolving in a hundred volumes of water at 34° F. Fish breathe oxygen. Were it not for this oxygen in solution, they would drown.

Oxides. At high temperatures oxygen is an exceedingly active element. A heated piece of iron burns in it, and most elements, as sulphur, phosphorus, carbon, copper, zinc, and magnesium, combine directly with it. These compounds of elements with oxygen are called *oxides,* and are very common in nature. Water might be called an oxide of hydrogen, while sand is an oxide of silicon.

Oxidation and oxidizing agents. The combining of oxygen with any substance is called *oxidation,* and, if the combining is accompanied by light and heat, it is called *combustion.* Substances such as potassium chlorate, that readily cause oxidation, are called *oxidizing agents.* If you throw a fragment of charcoal into some molten potassium chlorate, it will ignite and burn furiously. The charcoal is oxidized and the potassium chlorate is the oxidizing agent. It is the oxygen in the air that supports burning and life.

Oxygen is an element in which we should be intensely interested, for our existence depends upon it. It is the oxygen that we breathe in from the air that keeps us alive. Our fires burn because the fuel combines with the oxygen of the air. The organic waste of the world disappears because it is oxidized by oxygen. There is perhaps no other element that we use so continually.

Ozone. A silent electrical discharge is found to change the properties of oxygen in many ways. In the first place three volumes of oxygen contract to form two volumes of a new gas. The new form of oxygen thus produced is called *ozone.* It is a gas of irritating odor and is an active form of oxygen. It is a more powerful oxidizing agent than oxygen. For example, silver is not acted upon by ordinary oxygen at any temperature, while ozone attacks it, forming black silver oxide. Ozone cannot be kept for any considerable length of time, as it slowly turns back into ordinary oxygen. A number of other elements also exist in different forms. Carbon, for example, exists as the diamond, graphite or black lead, and lampblack. All these different forms consist only of particles of carbon, yet the physical properties of the different forms are dissimilar. Such forms are called *allotropic modifications.* Ozone is such an allotropic modification of oxygen.

Uses of ozone. The most important use of ozone is in the purification of drinking water. When used in the proper concentration, it will kill any number of bacteria, and attacks the pathogenic or dangerous germs first. The German Imperial Board of Health has shown that one gram of ozone will kill 30,000 cholera bacteria per cubic centimeter, in 250 gallons of water. It is much used in France and Russia to purify water.

Ozone does not kill the bacteria when used to ozonate air, but does remove the odors due to tobacco smoke, perspiration, etc. Owing to its strong oxidizing power, it is an ideal bleaching agent for oils, but its high cost prevents its commercial use for this. Owing to its extreme activity, ozone is a poison, and must be used with care.

Hydrogen peroxide. Like many other elements, hydrogen and oxygen combine in more than one proportion. In water, one part of hydrogen unites with eight parts of oxygen by weight; in hydrogen peroxide, one part of hydrogen unites with sixteen parts of oxygen by weight. That is, for every one part by weight of hydrogen, we may have either eight or sixteen parts by weight of oxygen. You will notice that these two numbers, eight and sixteen, are in the simple ratio of 1 : 2. The importance of this fact you will see later, when we come to study the theory of chemistry. This new compound, hydrogen peroxide, is not formed by the direct union of the elements.[1]

Hydrogen peroxide, when pure, is a thick, colorless sirup, which decomposes easily. It is, in fact, impossible to keep it for any great length of time, as it decomposes, giving off oxygen and leaving water. This makes it useful as an oxidizing agent. Whenever it is mixed with anything that can be easily oxidized, the oxygen combines with the other material and destroys it. Thus, if hydrogen peroxide is poured on the pus in a wound, the pus is oxidized, and the wound cleansed. In medicine it is used as a disinfectant because it leaves only water as a residue. A 3% solution of it in water is sold under various names, as hydrogen dioxide,

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[1] The commercial way of preparing hydrogen peroxide is to heat the oxide of a rare metal, barium, until it has taken up an extra portion of oxygen, forming barium peroxide. Barium peroxide is then mixed with cold, dilute sulphuric acid. The barium and hydrogen exchange places, forming barium sulphate and hydrogen peroxide.

hydrogen peroxide, or dioxygen. It is a safe and cheap home remedy, and should be more widely used than it is.

Bleaching power of hydrogen peroxide. Chemically, it has a second important use. There are many substances that need to be bleached, but the ordinary bleaching agents, as chlorine, are so powerful that they not only bleach, but attack the material itself. For these substances hydrogen peroxide is an excellent bleach, as it oxidizes the coloring matter without injuring the material itself. Silk, wool, and ivory are all bleached by its use.

SUMMARY

Oxygen is prepared by the electrolysis of water, or by heating a mixture of potassium chlorate and manganese dioxide. It is a colorless, odorless, tasteless gas. At high temperatures it is very active, combining with most elements to form oxides. It supports burning and life.

An oxide is a combination of oxygen with one other element.

Oxidation is the combining of some substance with oxygen.

An oxidizing agent is a substance that readily gives up oxygen, thus oxidizing other substances.

A catalytic agent is a substance that aids chemical change without itself being permanently changed.

Ozone is an active form of oxygen and is a good oxidizing agent.

Hydrogen peroxide is a good oxidizing agent, and because of this is a good bleaching agent and disinfectant.

Exercises

1. Why do goldfish, kept in a glass globe, die if the water is not often changed?

2. If oxygen is an active element, why does not the coal in coal mines burn up?

3. Name three common oxides that you will find in every house.

4. Why does the cork of a hydrogen peroxide bottle often fly out?

5. Why is the cork of a hydrogen peroxide bottle bleached?

CHAPTER V

HYDROGEN

Occurrence of hydrogen. Every organism is made up of many different compounds. One element, however, exists in practically all of these compounds, namely, hydrogen. It is found combined with carbon in almost every organic compound. Meat, kerosene, candles, fats, and oils all contain it. Hydrogen is found free in nature only in small quantities, but we have reason to believe that large quantities exist in the sun and stars. Natural gas, rock salt, and meteorites all contain it in small quantities.

Preparation of hydrogen. To prepare hydrogen we may decompose water by electricity, or by the action of certain metals (as sodium) on it; or, best of all, we may obtain it from an acid by the action of a metal upon it. Zinc, for example, is placed in a flask fitted with a thistle tube and a delivery tube (Fig. 13). Dilute sulphuric acid is added through the thistle tube. It acts upon the zinc, forming zinc sulphate and setting hydrogen free. The hydrogen is then collected by water displacement. Other metals, as iron, and other acids, as hydrochloric acid, may be used to prepare hydrogen.

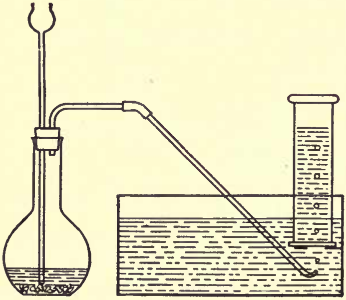

Fig. 13. — Preparation of hydrogen.

Properties of hydrogen. Hydrogen is a tasteless, odorless, colorless gas. It can be condensed to a liquid under high pressure at an extremely low temperature. It is the lightest gas known, air being 14.43 times as heavy as the same volume of hydrogen. It burns with an almost colorless flame, giving an intense heat. A pound of hydrogen, in burning, gives out more than four times as much heat as a pound of the best coal. It is therefore a most excellent fuel, but owing to its cost and bulk, it is not often used pure. Mixed with other gases, hydrogen forms illuminating gas, and those of you who have gas stoves at home know how convenient a fuel that is.

The lightness of hydrogen is made use of in balloons. As air weighs 14.43 times as much as hydrogen, volume for volume, a balloon filled with hydrogen will lift a considerable weight. The German military balloons are filled with hydrogen.

Oxyhydrogen blowpipe. When we wish an especially intense heat, such as is required to melt platinum, which is one of the exceedingly infusible metals, we resort to the *oxyhydrogen blowpipe* (Fig. 14). This is a device consisting of two concentric tubes. Through the outer larger one hydrogen is forced, and through the inner tube oxygen. These gases are both used under pressure, and give, when they combine, an intensely hot, pointed flame. This flame is sometimes used to produce an exceedingly bright light. It is itself almost invisible, but when it falls upon a piece of quicklime, which is infusible even at that high temperature,

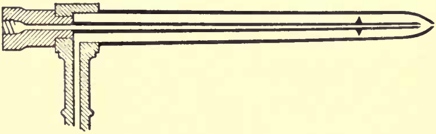

Fig. 14. — Oxyhydrogen blowpipe.

the lime is heated to incandescence and a dazzling bright light results. This is the so-called *calcium* or *lime light* which is used in theaters and magic lanterns (Fig. 15). It is being replaced, however, to a large extent by the electric light, which is brighter, cheaper, and more convenient.

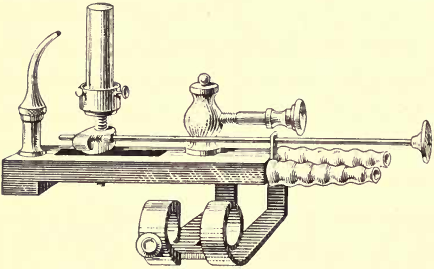
FIG. 15. — Burner for lime light.

Hydrogen a reducing agent. The action of hydrogen is, in many ways, the reverse of that of oxygen. If we heat oxide of copper in a stream of dry hydrogen, the hydrogen will combine with the oxygen, forming water, and leaving metallic copper (Fig. 16). Such an action as taking away oxygen from a compound is called *reduction*. It is evidently the reverse of oxidation. Hydrogen is then a *reducing agent*, and is

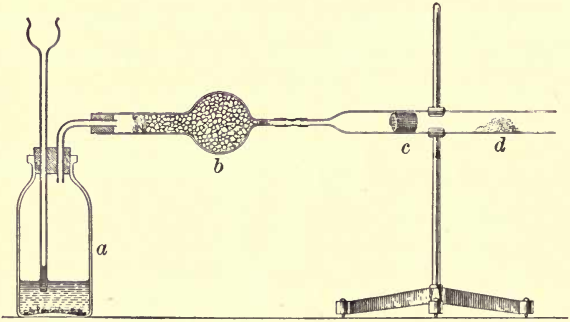

FIG. 16. — Reduction of copper oxide by hydrogen. *a*, hydrogen generator; *b*, calcium chloride for drying; *c*, copper oxide; *d*, anhydrous copper sulphate.

largely used in the laboratory for that purpose. Carbon is another excellent reducing agent, and is used commercially to extract metals from their ores, which are largely oxides. Iron is made in enormous quantities by this process.

SUMMARY

Hydrogen is prepared by the electrolysis of water, by the action of sodium on water, or by the action of a metal on an acid. It is a colorless, tasteless, odorless gas. It is the lightest gas known. It burns with an intense heat. It is a good reducing agent.

A reducing agent is a substance that readily takes oxygen away from its compounds.

The oxyhydrogen blowpipe is used to produce an intense heat.

Exercises

1. How can you prove that kerosene contains hydrogen?

2. Why is it very difficult to keep hydrogen, even in tightly corked bottles?

3. Do you regard a mixture of hydrogen and oxygen as an explosive? Explain.

4. One should never apply a flame to apparatus in which there is hydrogen without testing to see that the hydrogen is pure. It is not safe to guess. Why?

CHAPTER VI

ATOMIC THEORY

The alchemists. All through the Middle Ages, hundreds of men, called alchemists, worked in laboratories, studying what we now call chemistry. They discovered many chemical compounds, and found out many interesting and useful facts about them. They learned how to make sulphuric, hydrochloric, and nitric acids; how to extract some of the metals, as zinc and lead, from their ores; how to purify substances by distillation; in short, they knew and used many of the substances and operations that you have used in your work in chemistry. Yet they knew little or nothing of what we, to-day, call the science of chemistry. This was due to several reasons. They were not trying to find out the laws of nature, or to build a science, but to learn how to transmute base metals into gold, to make the elixir of life, and the philosopher's stone. They had no idea of the use of the balance, or any orderly way of working. They made their discoveries by mixing substances at random, and then seeing and recording what happened. Nor was one man willing to help others by publishing what he had found out. They learned a considerable number of isolated facts, but failed to discover those laws and quantitative relationships on which modern chemistry is based.

Chemistry a science. It was not until the time of the French Revolution that chemists realized that the science of chemistry must be based upon knowing not only what chemicals are used in chemical changes, but how much of each, the weight of the products formed, and, most important of all, the laws governing these changes. It was at this time that chemistry first really became a science, and that chemical facts were studied with the aim of discovering the laws governing them.

Law of Definite Proportions. Just as soon as chemists began not only to pull things apart, but to weigh the products obtained, a remarkable fact became known. It was found that the composition of every chemical compound was always the same. Thus, water was found always to contain 88.8^{+} % of oxygen, and 11.1^{+} % of hydrogen by weight. Or, 8 pounds of oxygen always combines with 1 pound of hydrogen, to produce 9 pounds of water.[1] No matter where the water comes from, if pure, its composition is found to be always the same. Hundreds of other compounds have been analyzed, but always with a similar result. If a substance is pure, its percentage composition is always the same; that is, a compound is always made up of the same elements combined in the same proportion by weight. This fact has come to be called the Law of Definite Proportions, and may be stated thus: *The composition of every pure chemical compound is always the same.*

Law of Multiple Proportions. Soon after the discovery of the Law of Definite Proportions, it was found that oxygen and hydrogen combined not only in the proportion of 8 to 1, but

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[1] It must be remembered that, for the sake of clearness in these notes, many facts are stated in round numbers. Thus oxygen and hydrogen really combine in the proportion of 8 of oxygen to 1.008 of hydrogen. The exact values are in most cases given in the tables in the appendix. In solving problems, the approximate values only are to be used.

that a second compound existed in which the proportion was 16 to 1. This second compound, hydrogen peroxide, you have already studied. You will notice that in these two compounds, the amounts of oxygen that combine with one pound of hydrogen, namely, 8 to 16, form the simple ratio with each other of 1 to 2.

Compounds of many other elements have been found that show this same fact. Carbon, for example, forms two oxides. We can combine 16 pounds of oxygen with 12 pounds of carbon, forming carbon monoxide, the poisonous gas found in our illuminating gas. Or we can combine 32 pounds of oxygen with 12 pounds of carbon, forming carbon dioxide, that gaseous substance that makes soda water effervescent. That is, 12 pounds of carbon will combine with either 16 or 32 pounds of oxygen. Here, again, you will notice that the amounts of oxygen that combine with the same amount of carbon form a simple ratio, 1 to 2.

The same fact is true of compounds of iron with sulphur, oxygen with nitrogen, and of the compounds of many other elements. From these facts we have formulated the following Law of Multiple Proportions: *If two elements, A and B, combine to form more than one compound, the weights of the element B that combine with a fixed weight of the element A, bear a simple ratio to each other.*

Matter is non-continuous. Our next task is to find an explanation as to why these laws are true. There are only two possibilities as to the constitution of matter. Either matter must be a continuous substance, or it must be made up of small particles held together by some force. Many facts show that the latter view is the correct one. To the eye, a plate of iron seems to be a continuous solid. We know that in reality it is not, for by applying an enormous pressure, it is possible to force water through the solid iron. The iron does not break, yet drops of water find their way through. This must be because the iron is made up of very small particles, and the water finds its way between them, showing not only that the iron is made up of small particles, but that these small particles do not touch each other. This gives us an explanation how it is possible for a bar of iron to expand when heated without undergoing any change in weight. The small particles that compose it have been driven further apart, and thus occupy more room. Cooling and hammering drives them closer together, and so the bar becomes smaller, although its weight does not change.

Matter is made up of small particles. If we mix a pint of alcohol with a pint of water, we naturally expect to have two pints of the mixture. Instead, we get only 1.8 pints. Evidently this is a case similar to the mixing of a pint of coffee beans and a pint of granulated sugar. There would not be a quart of the mixture, as the sugar sifts in between the larger coffee beans. In the same way, the alcohol and water being in reality made up of tiny particles, and these particles not filling all the space, the smaller particles sift in between the larger ones, thus reducing the volume of the mixture. There are just as many particles of water and alcohol as before, only the space between them has diminished and therefore the volume has become less. The weight of course remains unchanged. It is impossible to explain such facts as these except by believing that matter is made up of small particles, and that these small particles do not completely fill the space occupied by the body.

Molecule defined. A drop of water is made up of countless myriads of exceedingly minute particles. If you will in imagination divide and subdivide a drop of water, you must at last have a particle so small that, if you divide it again, you will no longer have water, but oxygen and hydrogen, the elements of which water is made. This smallest particle of water that can exist as water is called a *molecule.* We may define a *molecule* as the *smallest particle of any substance, that can exist in the free state and still be that substance.* Wood, sugar, salt, iron, all seemingly so solid, are in reality made up of molecules.

Molecules are so small that it is impossible to see them, even with the aid of the most powerful microscope. Small as these molecules are, it is possible to obtain by physical methods some idea of their size. Lord Kelvin has calculated that the distance between centers of contiguous molecules in a solid must be not more than $\frac{1}{250,000,000}$ inch, and may be only one half of this distance. If then you will cube 250,000,000, you will obtain a number that gives the smallest possible number of molecules in a cubic inch. The number is so huge that we are totally unable to obtain any real idea of its value. If you counted one molecule each second, worked ten hours a day, and 300 days a year, it would take you more than 23 years to count one line of molecules one inch long.

If you will imagine a drop of water, magnified until it is as large as the earth, the water molecules would be somewhat smaller than baseballs. It is difficult for us to imagine the existence of these molecules, yet it is only by believing in their existence that we can explain many of the laws of chemistry and physics.

Atoms defined. You have already learned that certain substances, as oxygen, gold, and iron, are called elements because we have found it impossible to pull them apart and produce from them other and simpler substances. You have also seen that by combining elements it is possible to produce new substances, as when you combined oxygen and hydrogen to produce water. We can prove that one molecule of oxygen will combine with two molecules of hydrogen to form two molecules of water. Since each molecule of the water produced contains oxygen, we must have split the oxygen molecule into two parts. The oxygen molecule must then be made up of two still smaller bodies, and to these we give the name of *atoms*. We may then define an atom as *the smallest part of the molecule of an element that can combine with other atoms to form molecules.*

The difference between the molecule of an element and the molecule of a compound is that the atoms in the molecule of the element are all alike, while the atoms in the molecule of a compound are different. The molecule of oxygen contains two small bodies, exactly alike, called atoms of oxygen, held together by a force that we call chemical attraction (Fig. 17). The molecule of water, on the other hand contains two like atoms of hydrogen and one of oxygen, different from the other two, all three atoms being held together by this force of chemical attraction.

FIG. 17. — 1. A molecule of oxygen.
2. A molecule of water.

Atoms of an element are all alike. The Law of Definite Proportions proves to us that all atoms of the same element are alike. If every atom of oxygen did not have exactly the same weight as every other atom of oxygen, then, in water, a compound containing oxygen atoms, the proportion by weight in which the oxygen occurs would sometimes vary. The same thing is true of atoms of all other elements. If we could divide a particle of any element minutely enough, we would obtain molecules first and then atoms. These atoms would all be exactly alike. Most of the common gaseous elements, as oxygen, hydrogen, and nitrogen, contain two atoms to the molecule.

Law of Multiple Proportions explained. How can we explain the Law of Multiple Proportions? Since the amounts of the second element present always form a simple ratio, we have evidently one, two, or three atoms of the second element present. That is, in the case of water, we have two atoms of hydrogen combining with one atom of oxygen. In the case of hydrogen peroxide, we have two atoms of hydrogen uniting with two atoms of oxygen. The weights of oxygen in the two compounds must be as one is to two, because the number of atoms is as one is to two.

Composition of molecules. Molecules of elements are composed of atoms, and these are all alike. Molecules of compounds are composed of atoms, and these atoms are different. Molecules of most elements contain two atoms, while molecules of compounds may contain from two to many hundred atoms. The molecule of albumen, the white of egg, contains over 250 atoms, while the number of atoms in a molecule of protoplasm is still greater. By ways that we cannot here discuss, chemists have determined the comparative weights of these tiny atoms, and can tell how many of them are present in a compound.

Symbols of atoms. Now that you know that matter is composed of atoms and molecules, you will be able to understand the way in which chemists write abbreviations of the names of these atoms and molecules. One molecule of sugar is composed of 12 atoms of carbon, 22 atoms of hydrogen, and 11 atoms of oxygen. To write all this, every time we wish to give the composition of a sugar molecule, would take too much time. Chemists have therefore agreed on a kind of shorthand, which is used by all chemists, no matter what their nationality. You may not be able to read German, but nevertheless you can, if you understand the simple principles used, read and understand the abbreviations in a German chemistry.

The symbol of one atom of any element is the first letter of its name, written as a capital. Thus, O means one atom of oxygen, C one atom of carbon. Where there are several elements, the names of which commence with the same letter, this system has to be somewhat modified. In these cases we use two letters, writing the first with a capital and the second with a small letter. For instance, one atom of carbon is C, one atom of chlorine Cl, and one atom of chromium Cr. Co is one atom of cobalt, but CO is one atom of carbon and one atom of oxygen.

In the case of some elements that have been known for a long time, we use in their abbreviations, or symbols, the Latin names. Thus, the symbol of one atom of sodium is Na, the abbreviation coming from the Latin name, natrium. Some others are: iron, Fe, from ferrum; copper, Cu, from cuprum; mercury, Hg, from hydrargeum. A complete list of symbols is given in the table of physical constants of common elements in the appendix (p. 378).

Formulas of molecules. Two atoms of hydrogen combine to form one molecule. We must have some way of writing the symbol of hydrogen so as to distinguish between two separate atoms of hydrogen and two atoms united to form one molecule. This we do by writing a figure before or after the symbol. H means one atom of hydrogen. It is not necessary to write the coefficient 1; it is understood. To indicate 2, 3, or more atoms of hydrogen, we write a coefficient in front of the symbol. Thus, when we wish to indicate two atoms of hydrogen, we write $2\,H$; for three atoms, $3\,H$.

When we wish to indicate that two atoms of hydrogen are combined to form a molecule, we write H_2. The symbol H_3 would be meaningless, as there are only two atoms in one molecule of hydrogen, and the symbol H_3 would mean a molecule made up of three atoms, which does not exist.

The formulas of molecules of compounds are written in the same way. A molecule of water is composed of two atoms of hydrogen, combined with one atom of oxygen. We might write the formula $2\,H\,1\,O$, but this would be inconvenient, as when we wish to write two molecules, we might confuse the coefficient 2 that meant two molecules with the coefficient 2 that meant two atoms of hydrogen. We, therefore, in molecules of compounds, write the number that expresses the number of atoms of each element present after, instead of before, the symbols of the element. We also write it below the line. The formula of water is, then, H_2O_1; but since it is unnecessary to write the 1, the formula is H_2O.

The formula of sugar is $C_{12}H_{22}O_{11}$. This means that each molecule of sugar is made up of 12 atoms of carbon, 22 atoms of hydrogen, and 11 atoms of oxygen. You will notice that the rule that there are two atoms of hydrogen in a molecule applies only to the element. In molecules of compounds there may be any number of hydrogen atoms present.

Physical and chemical change defined. Now that we know what matter is composed of, we can improve our definitions of physical and chemical changes. A *physical change* is one in which no new molecules are formed, but where the molecules are rearranged or separated. A *chemical change* is one in which new molecules are formed.

SUMMARY

Small units of matter. All matter is composed of atoms and molecules.

A molecule is the smallest particle of a substance that can exist and have the properties of the substance.

An atom is the smallest part of an element that can combine with other atoms to form molecules.

Law of Definite Proportions. The composition of every pure chemical compound is always the same.

Law of Multiple Proportions. If two elements, A and B, combine to form more than one compound, the weights of the element B that combine with a fixed weight of the element A, bear a simple ratio to each other.

Exercises

1. Is there any difference between $2\,H_2O$, and H_4O_2? Explain.

2. Is there any difference in meaning between $2\,H$ and H_2? Explain.

3. What difference would it make in the atomic theory, if we should find that lead was made up of two substances, X and Y, instead of being an element?

4. Name H, H_2, $2\,H$, $2\,H_2$.

5. Does H_3 have any meaning?

6. How many atoms are there in: $CuSO_4$, $KClO_3$, CO_2, HNO_3?

7. How many atoms are there in a molecule of cane sugar, $C_{12}H_{22}O_{11}$; vinegar, $HC_2H_3O_2$; washing soda, $Na_2CO_3 \cdot 10\,H_2O$?

CHAPTER VII

COMBUSTION

Combustion defined. In your work with oxygen you studied the burning of substances such as charcoal, sulphur, and phosphorus, and reached the conclusion that burning was a combination of the substance burned with oxygen. Chemists call " burning " *combustion*, and define it as any chemical change accompanied by light and heat. Oxygen is the great supporter of combustion, and the majority of cases of combustion that you will encounter are combinations of fuel, such as coal and wood, with oxygen.

Slow oxidation. The ordinary burning of wood is an oxidation accompanied by light and heat, called by chemists a combustion. Wood can, however, combine with oxygen so slowly that there is no evidence to our senses of either light or heat.

If you have ever wandered through a woodland, you must have wondered what becomes of the fallen trees. The ground is littered with fragments of bark and pieces of rotten wood, but despite the fact that trees have been growing in that ground for centuries, there is no large accumulation of wood from dead trees. We should expect to find huge piles of wood as a result of the death and overthrow of the trees, and the falling of twigs and boughs. But this dead wood has combined with the oxygen of the air. It has undergone a kind of burning. The same products, largely carbon dioxide and water, have been formed as if the wood had been burned in a stove. This oxidation, though, has been so slow that at no one time has heat enough been generated to raise the wood to a temperature at which it would give a large enough amount of heat to be perceptible to you. Such a slow combination with oxygen as this is called *slow oxidation.* Some chemists call it *slow combustion.* The final products are practically the same as in quick combustion, and the total amount of heat given off is the same, but the heat evolved is distributed over such a long time that the substance is not appreciably heated.

Kindling temperature defined. No substance will burn until it has been heated to a definite temperature called its *kindling temperature.* This temperature varies widely for different substances, but is always the same, for the same substance, under the same conditions. You have used one substance, phosphorus, which has a low kindling temperature. Iron, on the other hand, has a high kindling temperature, so high in fact that most people do not realize that it can be burned.

The way we build our fires well illustrates the use we make of various kindling temperatures. We first place paper, then wood, and lastly coal in the stove. Then we strike a match. Friction heats the composition on the head of the match to its kindling temperature, which is low, and the head burns. This heats the wood of the match to the temperature at which it ignites, and the match burns. We then touch the flame of the match to the paper in the stove. The paper, because its kindling temperature is fairly low, and because we need heat only a small amount of it, is easily raised to the required temperature and bursts into flame. The burning paper heats the wood to its kindling temperature, and this in turn heats the coal until it burns.

Kerosene burns freely, yet a lighted match can be thrust into a barrel of kerosene without danger. The match does not give enough heat to warm the kerosene to the temperature at which it takes fire. If, however, you should throw a cupful of kerosene on a blazing fire, it would burn fiercely. A large amount of coal placed on a feebly burning fire extinguishes it. There is not enough heat in the fire to warm the coal to its kindling temperature, and so the fire goes out. Had the coal been added gradually, each small portion would have been heated to its kindling temperature, would have caught fire, and the fire could have been saved.

Spontaneous combustion. There is one kind of combustion that is very dangerous. It is called *spontaneous combustion.* We know that when paint is first put on a house it is soft, but that in a few days it becomes hard. Paint is composed of linseed oil, mixed with some pigment, and the hardening of the paint is due to the oxidation of the oil, forming a solid substance. Many oils oxidize in the same way. As in every case of oxidation, heat is evolved. In the case of the paint on the house this heat passes into the air and does no harm. Suppose, however, you wipe up the kitchen floor with an oily rag, and then place the rag in a pasteboard box in a closet. The oil will begin to oxidize. This causes heat, and, as the heat cannot easily escape, the rag becomes warm. This in turn hastens the oxidation, more heat is evolved, and soon the rag is in flames.

Damp hay tightly packed in a barn, so that the air cannot circulate through it, will do the same thing. Even soft coal (bituminous coal) when it is in a fine powder will oxidize fast enough to catch fire, and ships have been lost at sea because of their cargo of coal catching fire from this cause. The remedy in the home is never to put away oily rags where the heat produced will be confined.

Some almost unbelievable things have happened as a result of unexpected oxidation. A manufacturer of small electric motors was in the habit of allowing the steel turnings he produced to accumulate until there were enough for a car load, when they were shipped to be remelted. One winter day, the watchman built a fire near the heap. The heat started the steel to burning, and for a whole day in spite of the streams of water that were played upon it, the heap of steel burned. Such things show that we cannot be too careful.

What to do in case of fire. The danger of being caught in a burning building is one to which we are all exposed, and, since in such an emergency most persons lose their self-possession, it will be well for you to think *in advance* what is the best course to pursue in such a case.

If awakened in the night by the smell of smoke, do not wait to dress completely, but wrap a blanket around you, and go out at once. Often the smoke is so thick that it is difficult to breathe. The air close to the floor will be found purer than that higher up, so you can usually get through a smoke-filled hall by crawling on your hands and knees. A wet cloth held to the nose and mouth is an aid, as it prevents the inhalation of the hot air and smoke.

Should your clothing catch fire at any time, *do not run.* This only serves to fan the flames. Instead, throw yourself on the floor and *slowly* roll over so as to smother the flames. Often a blazing skirt can be torn off, or the fire smothered by throwing a rug or coat over it. Above all, *keep calm.* If you do, you will escape with only a slight injury, while if you lose your head, you may lose your life.

Should a small fire start in your home, it can usually be extinguished by throwing water *on the burning substance,* not on the blaze. A bucket kept filled with water, and placed on the kitchen shelf, would often prevent a disastrous loss. It should have a *small* dipper kept in it, as it is better to throw the water on the fire in small quantities, rather than to pour the contents of the bucket on the fire at once. A siphon of vichy makes a very good fire extinguisher (Fig. 18). It is, however, useless to throw a small quantity of water on blazing oil or fat. Such fires should be smothered with sand, ashes, or flour, or a wet cloth or rug.

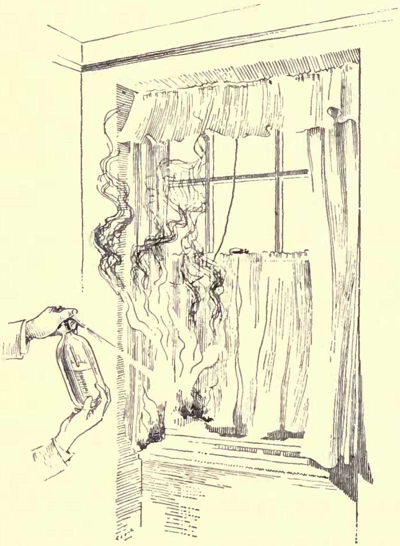

FIG. 18. — Extinguishing a fire with carbonated water.

If it is evident that the fire is beyond control, notify the fire department. You should know the position of the fire-box nearest your home, so that this notice can be quickly given. If you have a telephone in the house, use that. Here, again, the number of the fire house should be looked up in advance, and written on a card posted near the telephone. It is easy to do such things in advance, but most of us neglect them, with the consequence sometimes that our house burns down, when a little forethought would have saved it. Be careful to close the door when you leave the burning room to give the alarm. This tends to confine the fire to its starting point.

The average fire loss in the United States is over $200,000,000 annually. Much of this comes from carelessness in the use of matches. Never light a match in a dark closet filled with hanging clothes. Never leave matches where small children can play with them. Never keep matches in anything but metal boxes. Parlor matches are a source of danger, as mice gnaw the heads and thus cause fires. A match dropped on the floor may be stepped on, and so cause a conflagration. Safety matches cost but little more than parlor matches, and are much safer.

Methods of extinguishing fire. There are two general methods of extinguishing fires. We may lower the temperature of the burning substance below its kindling temperature, or we may smother the fire by shutting off the supply of oxygen.

Carbon tetrachloride, CCl_4, a volatile, non-combustible liquid, is a good fire extinguisher. A little thrown on a fire gives such a heavy vapor that air is excluded and the fire goes out.

Fireproofing fabrics. Conditions often arise, as in theaters, where light, combustible fabrics must be exposed to the danger of contact with flame. In such cases, it is desirable to fireproof the material. This may be done by the use of certain salts, as ammonium phosphate. Either of the following formulas may be used to fireproof cotton fabrics, by soaking them in the liquid, and then drying. A material treated with either of these will char, but will not burst into flame. The cotton bags in which flashlights are set off are treated with a similar solution. The fireproofing should be repeated every time the article is laundered.

(1)	Carbonate of ammonium	5 ounces
	Common alum	16 ounces
	Boracic acid	3 ounces
	Borax	3½ ounces
	Water	6 pints
(2)	Ammonium phosphate	8 ounces
	Boracic acid	1 ounce
	Water	4½ pints

A similar preparation is used on the wood from which safety matches are made. Wood, once ignited, glows for some time after the flame has been extinguished, and this spark is a source of danger. By impregnating the wood, this glowing is prevented, and a match carelessly thrown down cannot cause a fire.

If the articles are to be starched, the preparation below may be used.[1] The materials should be well mixed.

Hyposulphite of soda (granulated)	3 pounds
Cornstarch	3 pounds
Common salt	3 pounds
Borax (powdered)	1½ pounds
Talcum powder	3 pounds

The starch is made by taking four tablespoonfuls of this mixture and stirring it into three pints of boiling water. The mixture thickens, resembling ordinary starch, and is then ready for use, either hot or cold. The articles to be starched should be dipped into it, allowed to remain for one minute, removed, and wrung out. It is best not to remove too much of the mixture. Dry and iron as usual. It is not necessary to wax the iron, as it will not stick to the

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[1] This formula is due to Mr. A. J. Jarman.

cloth. The use of such a preparation would often avoid deplorable accidents.

Common fuels. There are three fuels, coal, kerosene, and illuminating gas, that we use so commonly that it will be worth our while to study them at this time. Coal, as you will find in a later chapter, is the product of the partial decay of vegetable matter. There are two principal varieties, anthracite, or hard coal, and bituminous, or soft coal. They both consist mainly of carbon, but soft coal contains in addition a considerable proportion of volatile matter. In the eastern cities, hard coal is used in the household, but elsewhere in the country soft coal is generally used.

When either variety burns, the carbon combines with the oxygen of the air to form carbon dioxide, CO_2, while any hydrogen present combines with oxygen, forming water. If, however, coal is burned using an amount of air insufficient for complete combustion, the carbon burns only halfway, and forms, instead of carbon dioxide, the gas carbon monoxide, CO. This gas is a poison, as it combines with the red coloring matter of the blood, and renders the blood incapable of absorbing oxygen.

How to regulate coal fires. Since ordinary burning is combining with oxygen, when we wish to burn coal in a stove or furnace, we must provide for a free access of air. In an ordinary range (Fig. 19) the air is drawn up through the grate bars, passes through the coal, burning it, and the waste gases formed then escape through the chimney. To control the fire, dampers are provided. These partially close the passage through which the gases escape, and so diminish the amount of oxygen fed to the fire. Thus the fire burns less briskly. Closing the doors under the fire does the same thing, that is, diminishes the supply of oxygen.

If the fire is not raked, the accumulation of ashes prevents the passage of the air, and the fire goes out. The ash that forms when coal is burned comes from the mineral matter in the vegetation from which the coal was formed.

It is often desirable to keep the fire over night when no heat is needed. This can be done by shutting the dampers, and lightly covering the fire with ashes. The fire will then smolder, and in the morning, a little fresh coal, raking, and opening the dampers will quickly revive it. Examine the stove or furnace in your house, and see if you can find out how the fire is controlled.

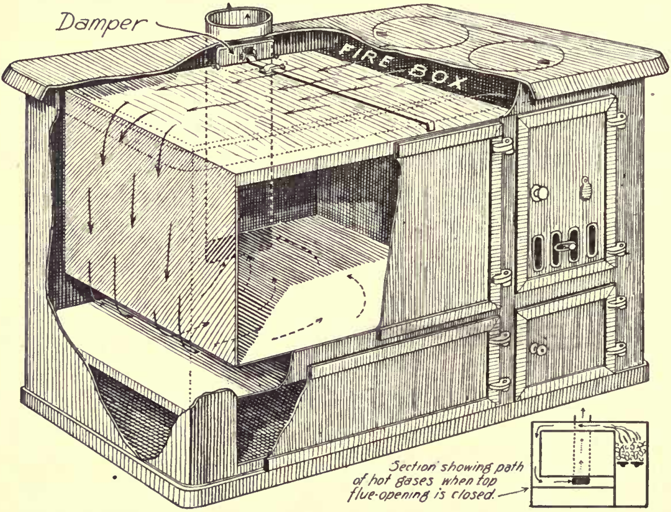

Fig. 19. — Diagram of cooking range.

Heating value of coal. Coal varies much in its heating power. One pound of good coal should be able to heat 14,000 pounds of water 1° F. Large corporations now buy coal on the basis of its fuel value, instead of by the ton. It would not pay you to have the necessary analysis made to enable you to do this, so you will have to trust to the coal dealer. One thing though you can do, and that is to watch the amount of ash that forms. Good coal should not give over 10% of ash by weight. If you find that your coal gives more than this, it is well to buy another brand next time. The veins of coal in the mine occur between layers of slate, and this is often not completely removed from the coal. You should, of course, not buy coal that on examination shows slate mixed with it.

Kerosene. Kerosene is a mixture of a number of compounds known collectively as hydrocarbons. They are all compounds of carbon and hydrogen, and on burning form carbon dioxide and water. If you have a kerosene lamp, notice that, when it is first lighted, a mist collects on the inside of the chimney. This is water formed by the burning hydrogen of the kerosene combining with the oxygen of the air. As the chimney becomes hot, the water evaporates. By keeping the chimney cold by surrounding it with a water jacket, the water will continue to be deposited until at last it will run down in drops.

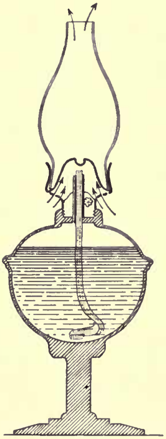

FIG. 20. — Lamp showing air currents.

From what you have learned about combustion you will see the advantage of the lamp chimney (Fig. 20). If we are to burn enough kerosene to give a bright light, we must have a plentiful supply of air. The chimney gives this by creating a good draft. If we turn the lamp too high, it smokes. Knowing the eagerness with which hydrogen and oxygen combine, you can explain this smoking. If we turn the wick too high, we vaporize so much kerosene that there is not enough oxygen in the air in the chimney to burn both the hydrogen and the carbon of the kerosene. Since oxygen would rather combine with hydrogen than with carbon, the hydrogen burns and leaves the carbon free. This carbon is the smoke that is sometimes so annoying, and blackens the ceiling over a lamp or a gas flame.

When we wish to manufacture lampblack, advantage is taken of this fact. We burn a hydrogen and carbon compound, as natural gas, in an insufficient amount of air. The flame smokes, and, on collecting the smoke, we have lampblack.

Luminous flames explained. To produce a flame, we must have a burning gas. In the case of the lamp, the heat of the flame converts the liquid kerosene into a gas. This burns, and gives the flame. You will remember that hydrogen burns with a flame, but that this flame is colorless. Since the only difference between the hydrogen and kerosene flame is the presence of carbon, the color of the flame must in some way be due to the carbon. The probable explanation is this. The heat of the flame decomposes the kerosene, and sets the carbon free in the form of very small particles. Since there is no oxygen in the interior of the flame, these carbon particles cannot burn, but are heated and glow with a yellow light. When they reach the edge of the flame, where they encounter the oxygen of the air, they burn, and, forming carbon dioxide, disappear. This explains why a thin, flat, wick is used. It gives more surface for oxygen to reach the carbon and so combustion is more perfect.

By lowering a cold plate over a kerosene or a luminous gas flame for a few seconds, you will find, on removing it, that carbon is deposited on the plate in the form of lampblack. This is because the cold plate cools the carbon in the flame below its kindling temperature, so that it cannot burn.

The Bunsen burner. Illuminating gas is a mixture of hydrogen, carbon monoxide, and some gases rich in carbon. The hydrogen and carbon monoxide burn and furnish heat, while the other gases set free carbon which, being heated, gives a luminous flame. The products of this combustion are of course the same as in the case of kerosene. In the gas stove, we do not need a luminous flame, but we do require a hot flame. We gain this by using what is called a Bunsen burner (Fig. 21). The difference between this and an ordinary burner is that, in the Bunsen burner, air is mixed with the gas before it is burned. Instead of the hydrogen and carbon having to pass through the flame, and come to the edge before burning, they burn in the flame. The carbon then does not pass in an incandescent state through the flame and so the flame is nonluminous. The Bunsen burner flame is much hotter than the ordinary luminous flame.

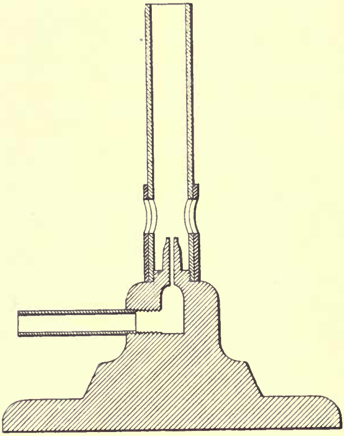

Fig. 21. — Section of a Bunsen burner.

To obtain a perfect Bunsen burner flame requires a nice adjustment of air and gas. This is accomplished by turning the regulator that opens and closes the air holes at the base of the burner, until the flame is colorless, but does not hiss. When using a gas stove, you will sometimes find that a kettle put over the flame becomes covered on the bottom with a coat of soot. This is because the burner is not properly adjusted. You can remedy the trouble easily, if you will examine the construction of the burner, and bear the above facts in mind.

A gas and air mixture is an explosive. A mixture of air and illuminating gas forms a violent explosive. Therefore, do not carry a flame into a room containing such a mixture. Many bad accidents have been caused by persons lighting a match to look for a bad gas leak. Bear this in mind when you light the gas oven. Some cooks have turned on the gas, and, after allowing it to mix with the air in the oven, have lighted it. You can imagine the consequences.

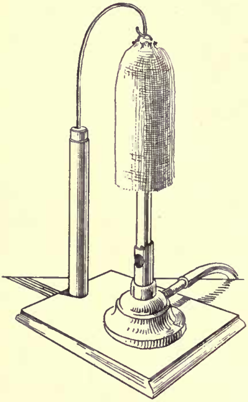

Fig. 22. — Gas mantle.

Welsbach gas burners. The intensely hot flame of the Bunsen burner, although it is itself non-luminous, may be used to produce a bright light. The oxides of certain rare metals, when heated to a high temperature, glow brightly. By immersing a loosely woven cotton wick in a solution of the mixed nitrates of these metals, drying it, and then burning the cotton, the nitrates decompose, forming oxides, which are left in the shape of the original wick. The result would be too delicate to transport. It is therefore dipped in collodion. This forms a flexible protective coating over the whole. In this form you buy it as the mantle of a Welsbach burner (Fig. 22). When the mantle is placed in position and set on fire, the collodion burns off, leaving the oxides behind as a fragile web. This is heated by the hot, non-luminous Bunsen burner flame of the Welsbach burner, and gives an intense light.

These burners use only a small quantity of gas, and are very bright. Their disadvantage is the delicate mantle. A slight jar is sufficient to break it to fragments. The inverted form of mantle is especially useful, as there is nothing under the light, and therefore the illumination is sent without waste where it is most useful.

Reading your gas meter. Stores are continually offering new forms of gas burners and gas heaters, and the statement is usually made that they burn such a small amount of gas that their cost is saved in a very short time. Naturally, the manufacturer of such devices is inclined to underestimate the gas consumption, and it is well for you to be able to find out for yourself just how expensive their use is. This you can do by learning to read the gas meter.

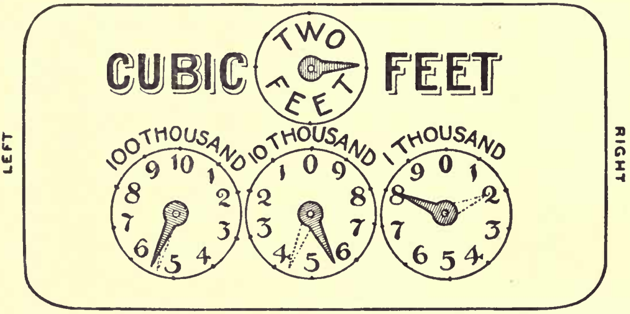

FIG. 23. — Gas meter dial.

When you study this lesson, as a part of your work, examine your gas meter. You will find a series of dials at the top, and over each dial a figure (Fig. 23). The arrangement is not always the same, but usually there are three dials, marked 1 thousand, 10 thousand, 100 thousand. Each dial is divided into ten parts, and these parts are numbered from one to ten. The number over the dial means the number of cubic feet of gas that one revolution of the pointer measures. Thus, when the 1 thousand dial hand has made one complete revolution, one thousand cubic feet of gas have passed through the meter.

The exact mechanical arrangement is difficult to explain without the use of an actual meter, but the principle is this. The meter contains a pair of bellows. The pressure of the gas causes the gas to flow into one of these bellows and fill it. As the bellows expands, it moves a wheel that is connected to the dial on top, and thus causes the pointer to go around. These pointers are so connected that one complete revolution of the 1 thousand hand advances the 10 thousand hand one tenth of a revolution. Thus it takes ten revolutions of the 1 thousand hand to make the 10 thousand hand go around once. As you burn the gas, it slowly flows out of the full bellows, which collapses, while at the same time the empty bellows fills.

The position of the pointers of these dials is read by the meter inspector, each month, and by subtracting the reading of the last month from the reading of this month, the gas company can tell how many thousand cubic feet you have used, and so what your bill will be. The next time your meter is read, go with the inspector and read it also. Tell him what your reading is, and ask him if it is correct. You will then be sure that you understand just how the amount of gas you use is determined.

At the extreme top of the meter there is usually a small dial that registers two cubic feet. By using this small dial you can tell just how much gas the various burners of your gas stove or gas lights use. Be sure that all the gas jets in the house are closed. Light the burner, the consumption of which you wish to know. Wait until the pointer on the small dial is exactly over a mark, and then note the exact time, using the second hand of a watch. When the hand has made one complete revolution, note the time again. You now know how many minutes that particular burner takes to burn two cubic feet of gas. A simple calculation will tell you the number of cubic feet burned per hour, and, knowing the cost of a thousand cubic feet, you can compute the expense per hour. In this way you can compare the efficiency of two styles of burners, or compute the expense of heating a room by the use of a small gas stove. It is more accurate if, instead of using only one revolution of the small hand, you allow the burner to use enough gas to cause the hand to make two or three revolutions.

This small hand may also be used to test for the presence of a leak. Note the exact position of the small hand in the morning, being careful to see that all the gas cocks in the house are turned off. Read its position in the evening, before any gas is turned on. A change in the position of the hand shows the presence of a leak.

Gas flatiron. One simple, useful gas device is the gas-heated flatiron. In the summer a small kitchen becomes uncomfortably hot, if the flatirons are heated over the gas stove on ironing day. This excessive heat may be avoided by using a gas iron (Fig. 24). It is like an ordinary flatiron, but is made hollow to allow the use of a Bunsen

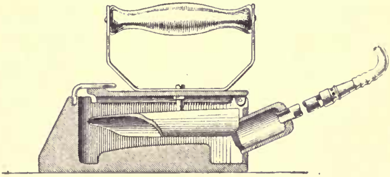

FIG. 24. — A gas-heated flatiron.

flame on the inside. This heats the iron, and avoids the necessity of changing irons. The consumption of gas is not large. If you have one at home, it will be worth while to test its consumption of gas and find out just what it costs to operate it. Test also the gas consumption of the burner that you would use to heat the ordinary iron, and you can tell whether it is a real economy to use a gas-heated flatiron.

How to use a gas stove economically. Care in the use of gas stoves will enable you to save a considerable proportion of your gas bill. In using a gas oven, for example, it is often possible, when the roast is almost cooked, to turn off the gas and let the heat of the oven finish the cooking. The use of a fireless cooker will still further cut the consumption of gas. You will find this described in another chapter.

SUMMARY

Combustion is any chemical change accompanied by light and heat.

Oxidation is union with oxygen. It may be slow or rapid. When it is rapid enough, it becomes combustion.

Kindling temperature is the temperature at which a body will begin to burn.

Flame is burning gas.

Kerosene forms water and carbon dioxide when it burns.

Luminous flames are often due to incandescent carbon.

A Welsbach burner light is due to the incandescent mantle.

Spontaneous combustion is combustion which occurs when no definite outside heat has been furnished to raise the body to its kindling temperature. It is brought about by slow oxidation under conditions which prevent the scattering of the heat so produced.

Exercises

1. What is an explosion?
2. Can flour dust explode? Explain.
3. Why is oily waste kept in metal cans in shops?

4. Can hydrogen be used in a Welsbach burner? Explain.

5. Why is gasoline not suitable for use in a lamp?

6. What would be the result of throwing a lighted match into a barrel of gasoline? Of sewing machine oil?

7. Why is water a poor thing to use to extinguish an oil fire? What would you use and why?

8. Why is a heap of fine soft coal dust dangerous?

9. Why is it necessary to rake the ashes from under the fire?

10. What does it mean, when your Welsbach mantle becomes covered with a black deposit? How would you remedy the trouble?

11. What should you do with a rag with which you have oiled the furniture? Why?

12. State in the order in which you would do them, all of the things you would do, if a fire broke out in your house.

13. Where does the fire escape on your house lead to?

14. Where is the fire alarm box nearest to your house?

15. How would you get to the roof of your house? Can you walk from the roof of your house to other houses?

16. If a fire should start in your house, would you go upstairs or downstairs, and why?

17. What steps should be taken to prevent an explosion in a mine full of fine coal dust?

18. Why are lace curtains and a gas flame, both near a window, a dangerous combination?

CHAPTER VIII

HEAT

Work defined. In the study of heat we need to understand the exact physical meaning of two words, *work* and *energy*. In physics, we do not regard everything that you might call labor as work. If you should support a pail of water in your hand, *without motion,* you would not be doing work in a mechanical sense. To do work on the pail of water, you must lift it. By *work we mean overcoming some resistance* (*as lifting a weight*) *through some distance.* The element of time does not come into the thought at all.

A convenient unit by which to measure work is the *foot pound.* If you lift one pound one foot, against the force of gravity,[1] you do one foot pound of work. It makes no difference how long you take to do this, the amount of work done remains the same. If you lift 6 pounds 4 feet, you do 6 × 4 = 24 foot pounds of work. You can measure work, then, by multiplying together the number of pounds lifted by the vertical distance through which they are lifted. It may interest you to calculate the foot pounds of work you do in going from the street to the top floor of the building.

Energy defined. *Energy is the capacity for doing work.* When you walk upstairs, you are lifting the weight of your body, or you are expending energy. Winding a watch gives energy to the spring. It is then capable of doing work in

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[1] We call the force that tends to pull bodies to the earth *gravity.*

making the wheels of the watch rotate. Coal possesses energy, for in burning, it may cause water to boil, and the steam thus produced may be used in a steam engine, which in turn may be used to lift a weight.

Our great source of energy is the sun. The energy contained in the sunlight causes the grass to grow. A cow eats the grass, and converts it into beef. You eat the beef, and convert its energy into the energy contained in the muscle of your arm. This muscular energy you use in winding up the spring of your watch. The energy thus given to the spring enables it to cause the wheels to rotate. This rotation causes heat, which passes off into space. In this series of energy transformations, you have changed its form, but have neither created nor destroyed energy. This fact, that we can neither create nor destroy energy, is called the *Law of the Conservation of Energy.*

Heat is a form of energy. For many years heat was thought to be a material substance called phlogiston. That is, a body, after it had been burned, was believed to weigh less than before, because the phlogiston had been taken from it. This belief we now know was incorrect, for bodies gain in weight when they are burned. This we can prove by weighing all of the products formed, the gases as well as the solids.

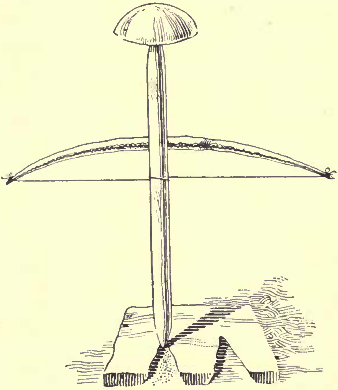
FIG. 25. — Primitive method of making fire by friction.

When two pieces of wood are rubbed together, they become hot. In fact, so much heat is thereby developed that savages have used this method to obtain fire (Fig. 25). Since the wood does not change in weight, mechanical motion (energy) must have been converted into heat. It is always true that a body in motion possesses energy, and that anything that arrests this motion converts this energy into heat. This energy change may be reversed. That is, heat may be used to cause motion, as when the energy contained in steam, due to its heat, is used to run a steam engine. Such facts as these cause us to believe that *heat is a form of energy.*

When we heat a body, the particles (molecules) composing it are set in motion. The faster the molecules vibrate, the hotter the body becomes. The energy that you put into the body is changed into molecular motion. The difference, then, between a cold and a hot piece of iron is simply that, in the hot piece the particles of iron composing it are moving more rapidly than are those in the cold piece. These iron particles, called molecules, are so small that we cannot see them, yet their velocity can be measured.

Fahrenheit and Centigrade thermometers. Calorie defined. You must carefully distinguish between degree of heat and amount of heat. Boiling water under standard (barometer reading 30 inches) conditions always has the same temperature, 212° F. A pint of boiling water, however, cannot contain as much heat as a quart of boiling water, although its temperature is the same. We measure degrees of heat by the aid of a thermometer. To measure quantity of heat, you must learn a new unit, the *calorie.* A calorie is the amount of heat that will raise the temperature of one gram of water 1.8° F. (1° C.).

There are two thermometer scales in use,— the Fahrenheit and the Centigrade scales (Fig. 26). The Centigrade scale divides the distance between the freezing and boiling points of water into 100 degrees, calling the freezing point of water 0° C., and the boiling point 100° C. On the Fahrenheit scale, this interval is divided into 212−32 or 180 degrees. Thus, 100 Centigrade degrees corresponds to 180 Fahrenheit degrees, or 1° C. = 1.8° F. The real definition of the calorie is, the amount of heat that will raise the temperature of one gram of water one degree Centigrade. Since we use the Fahrenheit scale in our work, we put instead of 1° C., its equivalent, 1.8° F. The large calorie, or Calorie, is equivalent to 1000 calories. In our study of foods we shall use *Calories* constantly.

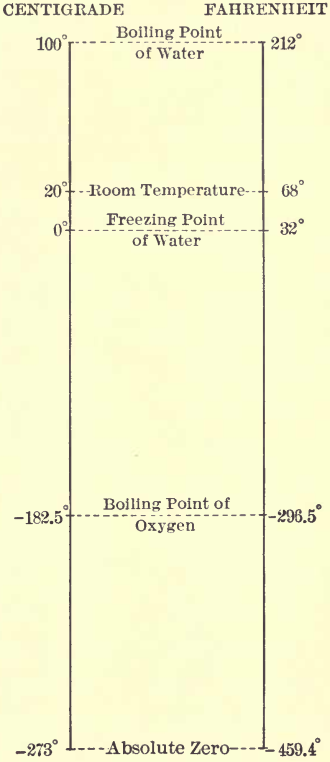

FIG. 26. — Comparison of Centigrade and Fahrenheit scale.

Absolute zero defined. Since the temperature of a body depends on the rate at which its molecules are vibrating, when a body is cooled, this molecular motion must become less. Evidently there must come a time when the molecules are at rest, and at this temperature there must be an absolute absence of heat. This temperature is called the *absolute zero*. On the Fahrenheit scale it is − 459.4°, on the Centigrade scale it is − 273°. Heating a body consists in causing its molecules to move more rapidly. This difference in the rate at which the molecules are moving is the only difference between a cold and a hot body. The hottest thing that we can make is the electric arc, which may reach a temperature of about 6000° F. The temperature of the sun is much higher than this.

Solids, liquids, and gases defined. It is the amount of motion of the molecules that determines the physical state of a substance, that is, whether it is a solid, a liquid, or a gas. In solids the molecules vibrate more or less rapidly, but do not change their relative positions in the body. The molecules attract each other. In liquids, the molecules not only vibrate back and forth, but are able to change their positions relative to each other. There is, however, a feeble attraction between the molecules. In gases there is not only no attraction between the molecules, but they repel each other, trying to separate as widely as possible.

These facts will enable you to understand why the following definitions are true. A *solid* is a substance that does not take the shape of the vessel in which it is placed. Solids have a definite weight, volume, and shape. A *liquid* is a substance that takes the shape of the vessel in which it is placed. Liquids have a definite weight, a definite volume, but no definite shape. A *gas* is a substance that takes the shape of any vessel in which it is placed, and distributes itself uniformly throughout the space. Gases have a definite weight, but neither a definite volume nor shape.

Sources of heat. The great source of heat in nature is the sun. If it should fail us, the earth would soon become a dead planet. Fortunately, there is no reason to anticipate any such calamity. Other sources of heat are the interior heat of the earth, heat caused by friction, and heat due to chemical action. In this last we are especially interested.

Effects of heat. When we heat or cool a body, that is, when we add to or subtract from its energy, a number of effects may be produced. The temperature of the body may change, or it may change its physical state, that is, a solid may melt, or a liquid be converted into a gas. Also the pressure of the body upon the containing vessel may change, as in gases, or its properties, as hardness, color, electrical conductivity, and volume, may vary.

Heat expands bodies. When a body is heated, it expands. This is true of solids, liquids, and gases, with practically no exceptions. The way in which water expands is somewhat unusual, and it will be well for you to refer back to the chapter on water and review the facts. The expansion of solids may cause the housekeeper expense and annoyance. If, in washing a large cut glass bowl, you put it at once into hot water, it is apt to crack. This is due to unequal expansion, caused by the outside of the bowl becoming hot, while the inside is still cool. The remedy is to put the bowl first into lukewarm water, and then to raise the temperature slowly by pouring in hot water. The same principle applies whenever breakable objects are to be heated. Suddenly cooling a hot glass will also crack it. Why?

The raising of cake is partly due to the expansion of a gas. When the batter is placed in the oven, the innumerable gas bubbles, derived from the reaction between the compounds of the baking powder distributed through it, are heated, the gas expands, and so the cake is made light. Why, then, do you think that banging the oven door, soon after the cake has been put in, is likely to make the cake fall? Those of you who have made popovers know how essential it is that the batter be very thin. If it is too thick, the force of the expanding gas is not sufficient to lift the upper crust, and a tough doughy mass results, instead of a light puffy one.

Heating bodies changes their physical state. Another effect of heat is change of state. If you leave an open pan of water exposed to the air, the water slowly disappears. It *evaporates.* This process will go on until either the pan is empty, or until the air has become saturated with water vapor. If you heat the water, you will find that the change from a liquid to a vapor proceeds more rapidly. When you have heated the water to 212° F., it boils. That is, bubbles of steam form all through the mass of the water.

This change of liquid to a vapor is called *vaporization* or *volatilization.* When the change occurs at low temperatures, the liquid is said to be *volatile.* Alcohol and gasoline are volatile liquids. Most solids when heated, first *liquefy* (melt), and then vaporize. A few, as sal ammoniac (ammonium chloride), iodine, and camphor, pass directly from the solid to the gaseous state and when this vapor is cooled, it returns to the solid state without passing through the liquid state. This we call *sublimation.* The substance is said to *sublime,* and the solid product obtained is called a *sublimate.*

Water will evaporate, even when it is in the form of ice. Wet clothes, hung out in the winter to dry, first freeze, and then dry by evaporation of the ice.

Fractional distillation. The rate of evaporation depends upon many things. First, upon the boiling point of the liquid. The nearer a liquid is to its boiling point, the faster it evaporates. Water boils at 212° F., while alcohol boils at 173° F. If, then, we heat both alcohol and water to 160° F., the alcohol will evaporate the faster, because it is nearer to its boiling point. Advantage is taken of this fact in *fractional distillation.* If we heat a mixture of alcohol and water, the alcohol will distill off first, leaving the water behind. The alcohol will not be quite anhydrous (water free), but nearly so. This process is often used to separate two liquids having different boiling points.

Conditions affecting evaporation. The rate of evaporation also depends on the extent of surface exposed. The larger the surface, the more freely the air will dissolve the liquid. Liquids to be evaporated should be placed in large shallow pans, rather than in deep narrow vessels. When the air over the liquid becomes saturated with its vapor, evaporation ceases. Evaporation in a deep vessel is very slow because the air is renewed with difficulty and quickly becomes saturated with the vapor. In chemical works, a blast of air is sometimes blown across the evaporating liquid, so as to bring dry air continually in contact with it. Would you expect clothes to dry faster on a windy or on a still day? Why? Do clothes dry faster on a cool or a warm day? Why? In the sun or the shade? Why? The drying of clothes is of much importance to the housekeeper.

Amount of water evaporated from the earth's surface. The total amount of water evaporated from the earth each year is enormous. The average yearly rainfall is between 30 and 40 inches. All this water (about 175 pounds for each *square foot* of the earth's surface) must first be evaporated by the air before it can fall as rain. The largest amount of this water comes from the ocean, but plants furnish more than is generally realized. In hot weather, grass loses by evaporation, each day, its own weight of water, or about 6½ tons per acre. Trees also contribute much water vapor to the air.

Boiling explained. If the temperature of a liquid is continuously raised, we at last reach a temperature at which the liquid boils. Bubbles of gas are then formed all through the mass of the liquid, and these gas bubbles in escaping set the mass of the liquid into violent agitation. The temperature then remains unchanged until the liquid has all boiled away.

The explanation of boiling is, that, as the temperature of the liquid is raised, its molecules are set into more and more violent motion. This motion increases until those molecules that are on the surface of the liquid are thrown so far into the air that they escape from the attraction of the liquid. We say that water boils at 212° F. This is not always true. From the explanation of boiling given above, you will see that the rapid escape of the molecules from the surface of a liquid must depend upon the pressure on that surface. If the pressure is diminished, the number of molecules present above the liquid is lessened, and it is then easier for new molecules to escape, or, in other words, the boiling point is lowered. If the pressure is increased, the reverse change takes place, or the boiling point is raised. This may be shown by half filling a round-bottomed flask with water, and boiling the water until the steam formed has driven all of the air from the flask. The flame is then removed, and the flask corked. If cold water is now poured on the flask, some of the steam is condensed (Fig. 27). This lowers the pressure, and the water inside the flask boils furiously. This may be continued until the water in the flask is so cool that you may place your hand on the flask, and yet the water inside is boiling.

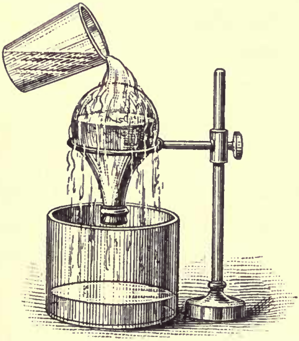

Fig. 27. — Boiling water at reduced pressure.

Advantage of boiling under diminished pressure. Many liquids cannot be boiled at ordinary pressures and temperatures without a chemical change taking place. When it is necessary to evaporate such liquids, we place them in a closed vessel, and pump off the air and vapor from above them. In this way the pressure on the liquid is diminished, the boiling point of the liquid lowered, and evaporation, without decomposition, becomes possible. This is always done in sugar refineries in evaporating sugar solutions.

Effect of altitude on the boiling point. As we climb mountains the air pressure becomes less. At the top of Mount Blanc, the pressure is so low that water boils at 183° F. In Denver, at an altitude of something over 5000 feet, water boils at 203° F. This makes a serious difference to a cook. The temperature of boiling water in Denver is so much lower than it is with us, that it is necessary to boil meat and vegetables there much longer than in New York. This diminished pressure does not of course affect the time required to roast meat, as, in roasting, the meat is exposed to the direct heat of the fire.

Effect of dissolved substances on the boiling point. The boiling point also depends upon the purity of the liquid. The effect of any dissolved solid is to raise the boiling point. Thus, water saturated with calcium chloride does not boil until a temperature of 354° F. has been reached. The presence of dissolved solids also lowers the freezing point of liquids. Sea water, for example, does not freeze at 32° F., but at a lower temperature.

Transferring heat by conduction. There are three ways in which heat is conveyed from place to place, *conduction, convection,* and *radiation.* A common case of the transfer of heat by *conduction* is in the use of a flatiron. The cold iron when placed on the stove is first heated on the bottom. The heating sets the bottom particles of the iron into rapid vibration, they strike against their neighbors, setting these in motion, and this continues until every particle of the iron is rapidly moving, or, until the entire mass of iron is hot. Iron is said to be a good conductor of heat, because heat easily travels through it in this way.

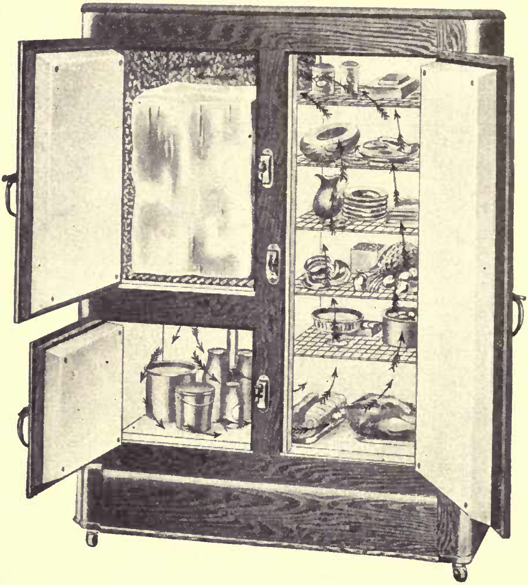

FIG. 28. — Circulation of air in a refrigerator.

Wood, on the other hand, is a poor conductor of heat. There is no difficulty in holding a lighted match, because heat travels so slowly through the mass of the wood, that, although one end of the match is burning, the other end is cold. Cloth is also a poor conductor of heat. It is for this reason that you use a cloth pad in handling a hot flatiron, as heat does not pass readily through the cloth. Metals, on the other hand, are good conductors of heat, silver being the best. You may have noticed how quickly a silver spoon, placed in hot coffee, becomes hot.

Air is one of the poorest conductors of heat. Woolen dress goods are warm for this reason. The wool fibers entangle large amounts of air among them, and it is this entrapped air that makes a woolen sweater warm. Linen, on the other hand, does not do this, and so linen is cool.

You may sometime when going skating have wrapped a newspaper around your body to keep you warm. It is really the air confined between the layers of paper that is the non-conductor, and so prevents the heat of the body from escaping. Do you think that a thick piece of cardboard would give you better protection?

A refrigerator is of necessity built of non-conducting materials (Fig. 28). An iron refrigerator would be useless, as heat would be so quickly conducted through it that ice in the interior of it would not keep for any length of time. In practice refrigerators are built of wood, with an inside lining of porcelain or metal. The space between these two walls is filled with various non-conducting materials (Fig. 29).

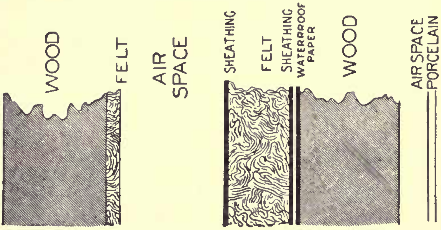

FIG. 29. — Section through the wall of a refrigerator.

Comparative conductivity of different materials. We can easily measure the comparative thermal conductivity of different substances. Roughly, if we represent the conductivity of silver by 10, the conductivity of copper is 8.5, iron 2, wood 0.003, and flannel 0.0004. These numbers are only approximations, but will serve to show the great differences that exist in the conductivity of different materials.

It will be an interesting home experiment for you to select two similarly shaped saucepans, one of enamel and one of steel, pour a glass of water in each, place them on the stove, and see in which case the water boils first. You will learn something about the conducting power of enamel that will be of use to you.

Convection. Heat is transmitted through air by a process called *convection.* The air over a hot surface is heated (Fig. 30), this air expands, and, becoming of a less *specific gravity,*[1] rises. Its place is taken by cooler air, and the process is repeated. By this means the air in the schoolroom is warmed.

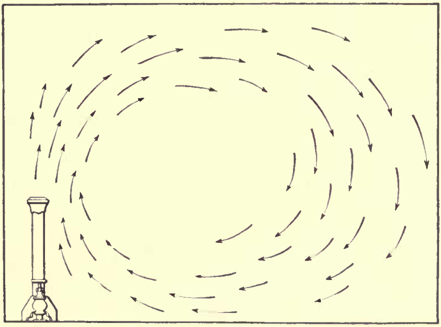

FIG. 30. — Circulation of air in room warmed by a radiator.

Since the warm air is continuously rising, the temperature is highest in the upper part of the room. Mothers sometimes forget this fact, and fancy, that, because

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[1] By *specific gravity* we mean the weight of a substance, compared to the weight of the same volume of some other substance that is used as a standard. For solids and liquids, this standard is water, for gases the standard is hydrogen. The specific gravity of sulphuric acid is 1.84, that is, any given volume of sulphuric acid (as one quart) will weigh 1.84 times as much as the same volume (one quart) of water.

the air is warm where they are sitting, it will also be warm enough for the baby playing on the floor. This is by no means always the case, and often the child suffers from cold, when the upper layers of the air in the room are warm enough.

Heat is transmitted in gases almost exclusively by convection. The same is largely true in liquids. Water is an extremely poor conductor of heat. If you will fill a test tube with water and, holding it by the bottom, place the top in a flame, you will find that you can boil the water in the upper part of the tube, while the water in the bottom of the tube is not even warmed.

The explanation of convection in liquids is the same as in the case of gases. The liquid in contact with the hot surface is heated, expands, becomes of less specific gravity, rises, and is replaced by the cooler liquid.

Radiation. *Radiation* is the third means of transfer of heat, and requires a little more detailed explanation. There is no air or any other form of ordinary matter between the earth and sun. It is therefore impossible that the heat of the sun should reach the earth by the processes of conduction or convection. We must imagine a totally different method of transmission.

We believe that all space is filled with an extremely rarefied substance that is called the *luminiferous ether*, or, more simply, the *ether*. (This is not the liquid used by surgeons to produce insensibility, but a totally different substance.) This ether is not like ordinary matter, but has some very wonderful properties. It is impossible to pump it out of any vessel. For, supposing it were possible to draw it out from the top, it would leak in through the sides, between the molecules of which the vessel was composed. Nor is it possible to pump more ether into a bottle than it already contains, for an attempt to do so would simply force ether out through the sides of the bottle. So far as we know, this ether can penetrate any body.

If we could magnify a drop of water sufficiently, it would somewhat resemble lemon jelly, filled, not too closely, with caraway seeds. The lemon jelly would represent the ether, and the caraway seeds the particles (molecules) of the water. Suppose now that you had a room filled with such a jelly, and should strike the jelly on one corner. A quiver would run through the entire mass of the jelly, just as it does in a mold of ordinary jelly on the table. This quiver may be called a *wave*. You know that when a wave moves through the water, the wave moves onward, but the water particles simply rise and fall as is shown by a piece of wood floating on the waves. So, in the case of the jelly, the wave moves through it, but the motion of any part of the jelly is only back and forth.

Imagine now the ether stretching from the sun to the earth. Some very hot particle at the surface of the sun, by reason of its high temperature, is in violent motion, and therefore has energy. This moving particle strikes the ether a blow and so causes in it a wave motion. This wave runs through the mass of the ether until finally it strikes the earth. There, the wave strikes some material, perhaps a stone. The energy of the wave is used up in setting the molecules of the stone in motion, that is, in heating it. That a wave can in this way transmit energy is shown by the destructive action of water waves in wearing away the shore.

You must understand that, while the wave is passing from the sun to the earth, it is *not heat*, but only a wave in the ether. It is not until this wave strikes some material object, that its energy is changed into heat.

The velocity of these ether waves is very great, about 186,000 miles per second. They can transmit not only heat energy, but light and electrical energy as well. It is by the use of ether waves that wireless telegraphy is possible. The difference between the ether waves that produce heat and those that produce light is in their wave length, that is, the distance from the top of one wave to the top of the next. Heat waves are longer than light waves.

You can now understand why a glass greenhouse is so much warmer than the outside air, even though it is not heated artificially. Ether waves coming from the sun can pass freely through the glass. These waves strike the plants and soil and are converted into heat. Glass is a poor conductor of heat, and so the heat cannot pass through the glass by conduction. The air in the greenhouse is confined, so that heat cannot escape by convection. The air in the greenhouse is therefore heated. The process of heat transfer by ether waves is called *radiation*, and the energy of such ether waves is called *radiant energy*.

Laws of radiant energy. Radiant energy obeys the same laws, whether we are dealing with light waves or heat waves. These waves are absorbed by dull, and reflected from polished, surfaces. They travel in straight lines. When they are reflected from polished surfaces the angle of incidence equals the angle of reflection, as is shown in Fig. 31.

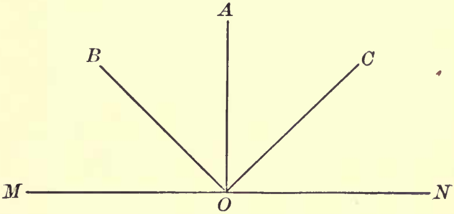

FIG. 31. — MN is the reflecting surface. $\angle BOA$ (incidence) $= \angle AOC$ (reflection).

These laws explain many facts that you are familiar with.

A white dress is cooler than a black one because radiant heat is reflected better from white than from black. Milk keeps better in a polished can than in a dull one for the same reason. The bottom of your teakettle should not be polished, because you want the heat to be absorbed and not reflected. It should not, however, be thickly covered with soot, for that is a poor conductor of heat. A screen placed in front of a fire protects you from the heat, because it reflects the ether waves. Many similar examples of the application of these laws to life will occur to you, if you will try to think of them.

Fireless cookers. The fireless cooker now so widely advertised is a practical application of a number of principles of heat. In boiling meat, for example, it is necessary to keep the meat for some time at about the temperature of boiling water. This can be done by heating it on the stove to the boiling point, and then placing the pot in some vessel provided with a non-conducting outside layer, such as a fireless cooker (Fig. 32).

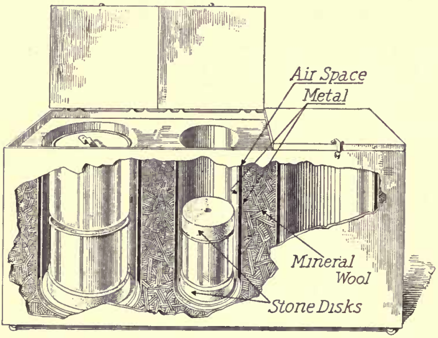

FIG. 32. — Fireless cooker.

A simple but efficient fireless cooker can be made at home. Fill a box with some non-conducting material as hay or sawdust. Leave a hole in the center in which to place the pot, and provide a cover, which also has a layer of cloth stuffed with hay over it. Such a simple device will keep a liquid hot for many hours. It may also be used as a refrigerator, for the non-conducting material will not only keep heat in, but will also keep heat out. Doubtless you can now tell why a woolen dress will keep you warm, and yet a piece of woolen cloth wrapped around a piece of ice will keep it from melting.

The manufacture of artificial ice. When a gas is liquefied heat is given off, and when the liquid changes back to a gas it absorbs heat. This principle is extensively utilized in the manufacture of artificial ice. Ammonia gas is cooled and compressed until it changes to a liquid. As a result of this compression, the liquid would become very hot, but it is cooled by keeping it surrounded by cool running water. This liquid ammonia is then run into a coil of large pipe, where it boils, changing back into a gas. The boiling point of liquid ammonia is very low, but it takes heat to change it into a gas just the same. This heat must come from the

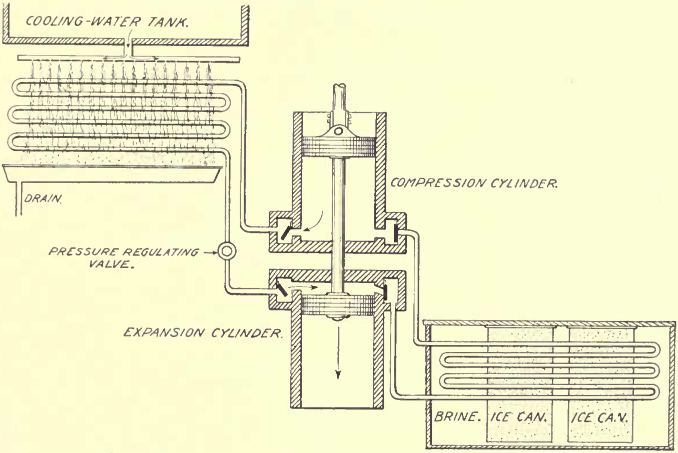

FIG. 33. — Manufacture of artificial ice by the ammonia process.

coil of pipe in which the ammonia is placed, so the pipe becomes very cold. This coil of pipe in turn is surrounded by a strong brine, and the cold pipe cools the brine. You have learned that dissolving solids in water lowers its freezing point, so, although the brine is cooled to about 16° F., it does not freeze. Cans containing pure water are suspended in the brine. This pure water is then cooled by the brine to 32° F., when it freezes. The ammonia gas that is formed is pumped out of the pipes, once more compressed to a liquid, and used over again.

Ice prepared in this way is purer than natural ice, and preferable for domestic use. One pound of it will do the same work in cooling that a pound of natural ice will; there is no difference, as both have the same temperature, 32° F. Ammonia is not the only substance that can be used in preparing artificial ice, but it has the advantages of cheapness and safety.

SUMMARY

Work is the overcoming of resistance through distance.

Energy is the capacity for doing work.

Law of Conservation of Energy: Energy can neither be created nor destroyed.

Heat is a form of energy.

Solids have a definite weight, a definite volume, and a definite shape.

Liquids have a definite weight, a definite volume, and no definite shape.

Gases have a definite weight, no definite volume, and no definite shape.

A calorie is the amount of heat needed to warm one gram of water one degree Centigrade. A Calorie equals 1000 calories.

On the Centigrade scale water boils at 100° and freezes at 0°.

On the Fahrenheit scale water boils at 212° and freezes at 32°.

Conversion of one scale reading to the other. To convert F.° to C.°:

$$(F.^\circ - 32) \div 1.8 = C.^\circ$$

To convert C.° to F.°:

$$(C.° \times 1.8) + 32 = F.°$$

Conduction is the transfer of heat from particle to particle.

Convection is the transfer of heat due to the mechanical motion of heated particles.

Radiation is the transfer of heat by radiant energy, or waves in the ether.

Exercises

1. Why is a fur coat so warm?

2. Should ice be placed in the top or bottom of a refrigerator? Explain.

3. Draw a diagram of a refrigerator showing the air currents inside.

4. We use electric fans in summer to cool ourselves, yet in winter an electric fan directed toward a radiator will cause a cool room to become warm quickly. Explain.

5. What is the principle of a Thermos bottle?

6. Glue is prepared from bones by heating them in water to a temperature of more than 212° F. How is this possible?

7. Water boils at 214° F., at the level of the Dead Sea. How can this be?

8. In warm climates, water is cooled by placing it in a porous jar (similar to a flower pot), and hanging the jar in the shade, but where the wind will strike it. Explain.

9. The temperature of islands is more equable than that of continents. This is due to the water surrounding them. Explain.

10. Dew seldom falls on a windy night. Why?

11. Why do double windows keep a room warm?

12. Should hot-air radiators be placed at the top or bottom of a room? Why?

13. Draw a diagram showing how your house is heated.

14. Why do not apples freeze when the temperature drops to just 32° F.?

15. Covering the ice in a refrigerator with a piece of carpet will cause the ice to melt very slowly. Is it well to do this?

16. How can food be cooked in a *fireless* cooker? Explain.

17. Why are water pipes buried deeply in the ground?

CHAPTER IX

THE ATMOSPHERE

Air is matter. That we live at the bottom of an ocean of an invisible gas, the air, is a fact that we quite generally ignore. It is only when this gas is set in violent motion, as in hurricanes, tornadoes, and cyclones, that we take notice of it, and yet our very existence depends upon its presence, for without oxygen, one of the gases found in it, we could not live.

We have found that steam is an invisible gas made by boiling water. If we could not cool steam and so change it to water, we would find it hard to believe that it was really composed of ordinary matter. Air, too, is such an invisible gas, and consists of ordinary matter. This is shown by the fact that it may be changed to a liquid by cooling and compression.

Since the boiling point of liquid air is below any temperature that naturally exists upon the earth (– 312° F.), air never exists in nature in the form of a liquid. When liquid air is cooled to a still lower temperature, it is changed to a transparent solid similar to ice.

That a so-called empty bottle is not really empty is easily shown (Fig. 34). If we tightly fit a cork, through which we have thrust a funnel, into the mouth of an empty bottle, and then pour water into the funnel, only a little water will run through the funnel into the bottle. The reason for this is that the bottle is filled with this invisible gas that we call air, and two bodies cannot occupy the same space at the same time. At first, a little water runs into the bottle. The air is compressed, since the cork prevents its escape, but soon the pressure of the water trying to get in is balanced by the pressure of the air trying to get out, and so no more water can enter. We can, however, by the aid of a second tube passed through the cork, let the air out, in which case the water quickly runs in and fills the bottle.

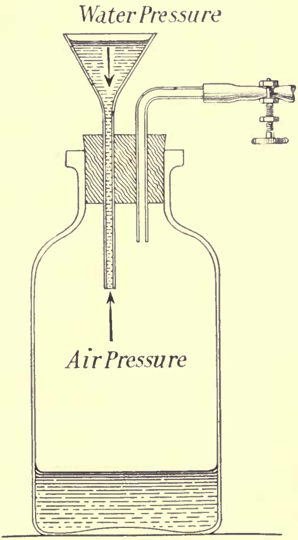

Fig. 34. — The so-called "empty bottle" contains air.

Air, then, in spite of the fact that we do not usually recognize its existence, occupies space, and is real matter. That air has weight can be shown by weighing a glass bulb fitted with a stopcock, and then pumping the air out of the bulb and weighing it again. The difference in the two weights must represent the weight of the volume of air in the bulb. If this volume is measured, the weight of a known volume of air can be determined. In this way the weight of a cubic yard of air has been found to be 2.18 pounds. That is, roughly speaking, 13 cubic feet of air weigh a pound.

It is this weight of the air, combined with its high velocity, that makes high winds so destructive, and that makes air offer such a high resistance to rapidly moving objects.

The barometer. The fact that air has weight can be shown in another way. Fill with mercury a long tube one square inch in cross section, and closed at one end, and invert it in a dish of mercury (Fig. 35). Since the end of the tube that dips in the mercury is open, you might expect the mercury in the tube to run out, and it does do so until the height of the mercury in the tube is about 30 inches. Then it remains stationary.

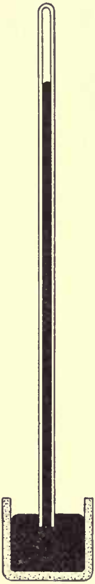

FIG. 35. — A Torricellian tube.

Since the mercury stays in the tube, something must be holding it there, and this something is the weight of the air pressing on the mercury in the dish. If we remove this pressure by placing the tube and dish under the receiver of an air pump and pumping out the air, the mercury will sink in the tube.

The mercury held up in the tube by the air weighs about 15 pounds. The mercury therefore presses down in the tube with a force of 15 pounds per square inch. Since the mercury is held up in the tube by the downward pressure of air on the mercury in the dish, which pressure is transmitted through the mercury in the dish to the base of the mercury in the tube, this pressure of the air must amount to 15 pounds per square inch. That is, the air presses down upon each square inch of the earth's surface with a pressure of 15 pounds. Or, in other words, the weight of a column of air one square inch in cross section and reaching from the earth as high as the atmosphere extends is 15 pounds. This pressure of 15 pounds to the square inch is often spoken of as a pressure of one atmosphere. This figure is only approximate, varying with the amount of water vapor in the air, and with the altitude of the place where the measurement is made.

This experiment with the tube was first tried by Torricelli, and the vacuum above the mercury in the tube is therefore often called a Torricellian vacuum.

Still another way to show this atmospheric pressure is to place a hollow tube, as a straw, in a glass of soda water, and then remove with the mouth the air from the upper end of the tube. The fluid is then forced up the tube by the outside pressure. This is commonly called "suction," but in reality the soda is pushed up the tube by the pressure of the air on the surface of the soda.

An inverted tube filled with mercury registers the pressure of the air, and as this pressure changes, the height of the mercury changes. Such a device is called a *barometer* (Fig. 36). By means of a barometer one can readily determine the daily changes in atmospheric pressure, due to temperature changes and the presence of water vapor in varying amounts.

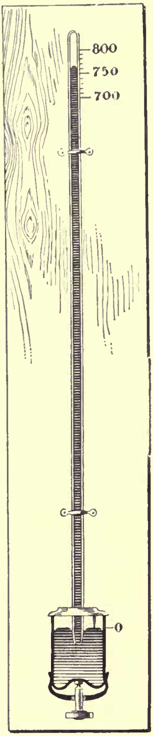

Fig. 36. — A barometer.

Use of barometer in foretelling weather changes. Water vapor is lighter than air. Since water vapor in the air, by crowding out the heavier gases, makes the air lighter, and thereby makes its pressure less, the mercury falling in the barometer tube (called a falling barometer) indicates an increasing amount of water vapor in the air, and hence a probability of rain. Thus changes of air pressure are frequently accompanied by the changes of weather with which we are familiar.

Generally, in an area where the air pressure is much lower than the average for that region, the weather is warm and cloudy; while in areas where the pressure is higher than the average, the weather is apt to be cold and clear. What we ordinarily call a storm is simply an area of low pressure and rainy or snowy weather.

It has been found by experienced observers in the United States that these related weather conditions of pressure, temperature, and cloudiness do not long remain in one place, but move across the country in a general direction from west to east. By a careful study of the usual paths and velocity of these areas, it has become possible to foretell the approach of a storm. The government publishes each day a map of the country with these conditions marked on it, and by its use it is possible to trace the path of a storm, day by

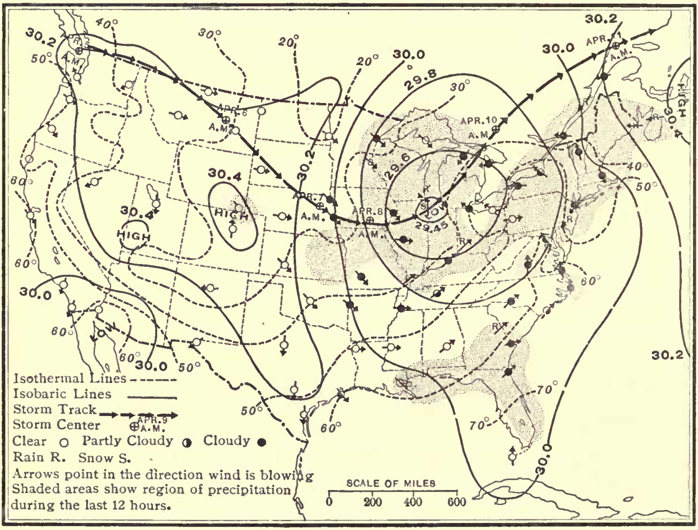

Fig. 37. — Path of a storm across the United States.

day, and to estimate when it may be expected to reach us (Fig. 37).

Absolute and relative humidity. Air always contains water vapor resulting from evaporation. The weight of water vapor actually present in a given volume of air is called its *absolute humidity.* This is usually expressed in grains per cubic foot.

The ability of air to dissolve water increases with the rise of its temperature. When air contains all the water vapor it can hold at a given temperature, it is said to be *saturated.* The ratio of the absolute humidity to the amount necessary to saturate the air at a given temperature is called the *relative humidity.* It is measured in per cent.

Suppose we find by actual experiment that there are 3.99 grains of water present in one cubic foot of the air in the laboratory. This would be the absolute humidity. We find the temperature to be 70° F. Now from a table (see Laboratory Manual) we find that, at 70° F., 7.98 grains of water per cubic foot are required to saturate the air. If 7.98 grains of water are required to saturate the air, and we have but 3.99 grains present, the air is $3.99 \div 7.98 = \frac{1}{2}$ saturated. The air is 50% saturated, or, the relative humidity is 50%.

Dew point, dew, and frost. If the temperature of the air is lowered, its capacity for holding water vapor is lessened, and therefore its relative humidity is raised. If the cooling is sufficient, the relative humidity will reach 100% (*saturated*), and further cooling will result in condensation of a part of the vapor. In this way *dew* and *frost* are formed.

The temperature when condensation of water vapor from the air begins is called the *dew point.* This is the cause of the " sweating " of ice-water pitchers, and the condensation of water on the window panes in the winter. Shopkeepers are often annoyed by this condensation of water on the glass of their show windows, and to prevent it, they direct blasts of air from an electric fan against the glass. This keeps it clear. Why?

Effect of humidity on our comfort. A good complexion is one of the things that we all admire, and a good complexion depends largely upon the humidity of the air. English girls are famous for the beauty of their skin, and England is famous too for its fogs and humidity. The connection is this. If the relative humidity sinks too low, water is taken from the skin by the air, and the result is a dry parched skin. If you travel in those parts of our country where the air is noted for its dryness, you will find that the skin of the natives presents a dry, leathery appearance.

When the humidity is low, not only the skin, but the mucous membranes give up water to the air, and we suffer in consequence. In winter, you know how often you go home and after being in the house a little while, begin to complain of a dry throat and a general feeling of discomfort. Often this is laid to steam heat. Steam heat is no different from any other kind of heat, the difficulty being that the relative humidity is 50% out of doors, and that it is a cold day. This cold air is warmed in the house to 70 ° F. but no additional water is added to it. The relative humidity will of course become less, although the absolute humidity remains the same. In consequence, the dry air takes moisture from our nostrils and throats, and we soon feel the effects.

The remedy naturally is to add water to the air. If you use steam heat, hang a wet towel on the radiator, and notice how quickly the room feels more comfortable. Furnaces commonly have a water pan provided, so that the warm air passes over water on its way to warm the rooms. This is right in principle, but the pans are usually made so small that they do not help much. Not only ourselves, but even the furniture feels the effect of this excessive dryness of the air. The glue that holds our chairs becomes so dried out that it cracks, and soon the chair falls to pieces.

An example may make this matter clearer. Suppose that, when the temperature outside is 20° F., the absolute humidity is 1.235 grains of water per cubic foot, or the relative humidity is 100%. If now we warm this air to 70° F., the amount of water vapor remains unchanged, but since at 70° F. a cubic foot of air can dissolve 7.980 grains of water, the relative humidity has fallen to 13%, an amount so low that we suffer. The relative humidity in our houses should be from 50% to 65% for comfort and health.

One advantage of a dry climate is the comfort that is experienced on a hot day. You have no doubt on some summer days felt hot and sticky, even though the temperature was not abnormally high. This is due to the fact that if the relative humidity is very high, the perspiration does not readily evaporate. That heat is required to evaporate liquids rapidly is a matter of common knowledge. If the relative humidity is low, the perspiration quickly evaporates. This requires heat, and this heat comes from our bodies, and thus we are cooled. In Arizona, with the thermometer at 100° F. you will be more comfortable than in New York at 90° F., owing to the dryness of the air in Arizona.

Preparing nitrogen from the air. If we float a piece of phosphorus, placed on a crucible cover supported on a cork, in a dish of water, set fire to the phosphorus, and then invert a bottle over it, the phosphorus burns for a while and then goes out (Fig. 38). A white solid, an oxide of phosphorus, results, and this solid being soluble in water, dissolves, and water then rises inside the bottle.

When we measure the amount of gas remaining, we find that about one fifth of the air has disappeared. One of the constituents of the air has been burned out. From your experiment on oxygen you will infer that the gas removed is oxygen. The gas remaining behaves in an entirely different manner from air. It is inert, will not support burning, we cannot live in it, nor will metals rust in it.

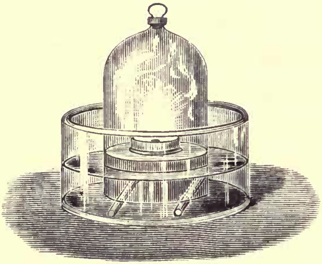

Fig. 38. — Removal of oxygen from air.

When we burn phosphorus in the air, we cannot easily regain the oxygen from the compound formed. If, however, we heat mercury in the air, it combines with the oxygen and changes to a red rust. If this rust is heated to a higher temperature, it breaks up into mercury and oxygen, thus giving us a means of separating air into its two main constituent gases, oxygen and nitrogen.

Composition of the air. The inert gas remaining after the oxygen has been removed is called nitrogen. It is important, because it is one of the constituents of an important class of foods. The composition of the air is about one fifth oxygen and four fifths nitrogen. The air also contains about four parts per 10,000 of carbon dioxide, a varying amount of water vapor, and small amounts of ammonia and certain rare gases, as argon and neon.

Proofs that air is not a compound, but a mixture. The oxygen and nitrogen in air are not chemically combined.

That the air is a mixture is shown in many ways, some of which are given below.

When air dissolves in water, if it were a compound, it would dissolve as air, that is, the proportion of oxygen in the dissolved gas would be the same as in air. Instead, we find that dissolved air contains twice as great a percentage of oxygen as the original air. This is because the oxygen and nitrogen dissolve in the water as separate gases, and not as a compound.

When pure water boils, the resulting steam, if condensed, has the same composition as the original water. If air were a compound, it would do the same, but when we liquefy air, and then allow it to boil, the nitrogen boils away first, leaving the oxygen.

The composition of air varies somewhat, while the composition of any compound is always the same.

Air essential to man. Air is just as essential to man as food. He absorbs daily, through his lungs, about 26 ounces of oxygen from the air, and exhales about 31.5 ounces of carbon dioxide during the same time. The inhaled oxygen dissolves in the red coloring matter of the blood (hæmoglobin), entering into a loose chemical combination with it, and is so carried to the cells of the body, where it is used in the oxidation of foods and protoplasm, thus making possible the development of heat and other forms of energy. Thus every movement of the body, the maintenance of the temperature of the body, and even our power to think depend on the oxygen we obtain from the air.

The carbon dioxide which results from this oxidation in the cell is carried back to the lungs by the blood and exhaled, consequently the air that leaves the lungs is of a different composition from that inhaled. The amount of oxygen has been decreased from 21% to about 16% by volume, the percentage of carbon dioxide has increased from 4 to about 430 parts per 10,000 in volume, and a large amount of water vapor has been added, as well as small amounts of various organic waste materials.

Ventilation. Exhaled air is generally warmer and therefore lighter than inhaled air, and therefore rises. In order to ventilate our houses properly, we should see that fresh air is admitted to the bottom of the room, and that the impure hot upper air is drawn off. It is therefore better, when we wish to ventilate a room, to open both the top and the bottom of the window (Fig. 39). A capacity of one thousand cubic feet of air per person in an ordinary room, with the usual ventilation through doors, chimneys, and cracks around the windows, will ordinarily keep the air in good condition.

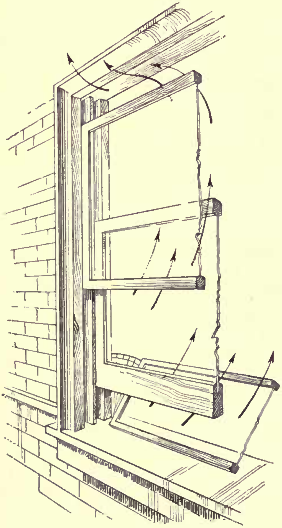

Fig. 39. — Ventilation by a device arranged to prevent a direct draft.

The carbon dioxide in pure air is about 4 parts in 10,000. In a room such as a school recitation room, it may rise to 20 parts, but this is evidence of serious organic contamination and the air is unfit to breathe. The New York Board of Education allows, in designing school buildings, 30 cubic feet of air per minute per pupil.

The notion that many people have that night air is injurious is unfounded. Circumstances may make some particular night air impure, but in general the air at night is just the same as during the day, and ventilation at night is just as important as ventilation during the day. To sleep in a strong draft may not be advisable, but to sleep in a room without good ventilation is to wake the next morning feeling dull, depressed, and out of sorts with the world.

What is known in history as the " Black Hole of Calcutta " is an illustration of what the lack of ventilation may cause. During the mutiny in India, 146 English prisoners were confined in a room 20 feet square, with only two small windows, which were obstructed by a veranda. As a result the next morning only 23 were alive. The others had been suffocated by the lack of oxygen. The same thing happens in a less degree when we live and sleep in improperly ventilated rooms. It is a slow but sure method of poisoning, since pure air is essential for our well-being.

The air cycle. Since not only animals but plants require oxygen and exhale carbon dioxide, and since every fire uses up the oxygen of the air, we may well wonder why it has not long since become exhausted (Fig. 40). The explanation is found in the way that plants grow. Almost every compound found in the vegetable world contains carbon, and this carbon comes largely from the carbon dioxide of the air. Plants take in the carbon dioxide exhaled by animals, and that produced by combustion, and use the carbon in forming wood and other com-

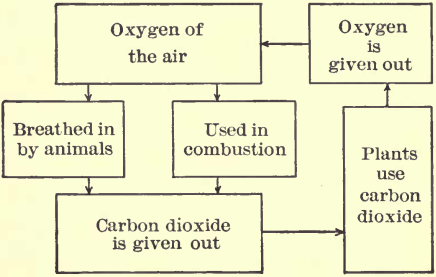

Fig. 40. — Cycle of carbon and oxygen.

pounds. The oxygen set free in this process passes off once more into the air. In this way the amount of oxygen in the air is maintained almost constant.

SUMMARY

Air can be cooled and compressed to a colorless liquid.

A pressure of one atmosphere is a pressure of 15 pounds to the square inch.

The weight of 13 cubic feet of air is about one pound.

The barometer at sea level usually stands at about 30 inches.

A falling barometer indicates the approach of a storm.

A rising barometer indicates clearing weather.

Absolute humidity is the amount of water vapor present in the air. It is usually expressed in grains per cubic foot.

Relative humidity is the ratio between the amount of water vapor actually present in the air to the amount required to saturate the air at that temperature.

The composition of the air is about $\frac{1}{5}$ oxygen and $\frac{4}{5}$ nitrogen.

Air is a mixture, not a compound.

Ventilation is essential to health.

Exercises

1. Carbon dioxide is harmless. Why then is a large percentage in the air of rooms regarded as objectionable?

2. Show by a diagram how you would ventilate your sleeping room.

3. Would it be an advantage to us to have the air pure oxygen? Explain.

4. There is little vegetation in a large city, and the city uses large amounts of oxygen from the air. Why do the citizens not die as a result of the exhaustion of the oxygen of the air?

5. Could plants live, if there were no animals?

6. Could animals live, if there were no plants?

7. Name two sources of the carbon dioxide in the air. Of the moisture in the air.

8. How is the composition of the air kept constant?

CHAPTER X

FORMULAS, EQUATIONS, AND VALENCE

Atomic weight defined. Every atom of an element is just like every other atom of the same element (p. 53). Every atom of hydrogen weighs just the same as every other atom of hydrogen. The symbol H then means not only one atom of hydrogen, but also a definite weight of hydrogen. Chemists are able to determine the comparative weights of atoms of different elements. The oxygen atom, for example, weighs 16 times as much as the hydrogen atom. The atom of carbon is 12 times as heavy as the hydrogen atom. These comparative weights we call *atomic weights.* (See table, p. 378.)

As the hydrogen atom is the lightest one known, its weight may be called one, and the weight of the other atoms expressed in terms of it. To say that the atomic weight of chlorine is 35.5 means that one atom of chlorine weighs 35.5 times as much as one atom of hydrogen. *Atomic weight* may be defined as *the weight of one atom of any element compared to the weight of one atom of hydrogen.*[1]

Molecular weight defined. The formula of water, H_2O, shows that water is made up of two atoms of hydrogen and one atom of oxygen. If we know the atomic weight of hydrogen to be one, and that of oxygen to be sixteen, we can

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[1] Since in determining these comparative atomic weights chemists usually work with oxygen, oxygen (16) is often taken as the practical standard.

calculate that a molecule of water weighs 18 times as much as an atom of hydrogen, the weight of which, you will remember, is our standard. This number, 18, is called the molecular weight of water. *Molecular weight* may be defined as *the weight of one molecule of a compound as compared with the weight of one atom of hydrogen.* It can always be found by adding together the weights of the atoms that compose a molecule. Sugar has the formula $C_{12}H_{22}O_{11}$. The 12 atoms of carbon weigh $12 \times 12 = 144$, the 22 atoms of hydrogen, 22, and the 11 atoms of oxygen, $11 \times 16 = 176$, making a total of 342 for the molecular weight of cane sugar.

Naming compounds. To learn how to name the thousands of inorganic chemical compounds would at first seem like a difficult task, yet by the aid of a few simple rules the difficulty vanishes. First, we must remember that in chemical formulas the symbol of the more metallic element is usually written first. When, then, we examine such a formula as KCl, we may not know what element the symbol K stands for, but we may be reasonably sure that it is the symbol of the more metallic element.

Naming binary compounds. The simplest compounds are of course those made up of only two elements, such as copper chloride, $CuCl_2$, potassium chloride, KCl, etc. Such compounds as these are easily named. All that you have to do is to name the more metallic element first, and then follow it by the name of the less metallic element, so modified as to end in *-ide.* In the formula $CuCl_2$ we know that the symbol Cu means copper, and since this is given first it is the metallic part of the compound. Cl means chlorine, and this name we change so as to end in *-ide.* The name of the compound is therefore copper chloride. Similarly, NaCl is sodium chloride.

The fact that the name of a compound ends in *-ide* tells you that it must be made up of two elements, and two elements only. Sodium oxide, for example, must be made up of sodium and oxygen; potassium iodide, of potassium and iodine.

It often happens that elements combine in more than one proportion. Thus, hydrogen and oxygen combine to form the compounds H_2O and H_2O_2. Both of these compounds may be called oxide of hydrogen. To distinguish between them the prefix *per-*, meaning the higher state of oxidation, is used in naming the H_2O_2.

Sometimes the Greek prefixes, *mono-*, one; *di-*, two; *tri-*, three, etc., are used to distinguish between different compounds of the same elements, as CO, carbon monoxide, and CO_2, carbon dioxide.

Naming chemical compounds. We can name most other inorganic compounds if we know the name of the acid from which they are prepared. All acids contain hydrogen, and the symbol of the hydrogen is always given first in writing their formulas. Nitric acid is HNO_3 and sulphuric acid, H_2SO_4. In many chemical changes involving acids, the hydrogen of the acid is replaced by a metal. In this way a new compound called a *salt* is formed.

When sodium and nitric acid interact, the hydrogen of the acid is replaced by sodium, and a new compound, $NaNO_3$, is obtained. Since this compound was made from nitric acid, and also contains sodium, its name should suggest both of these. It is called sodium nitrate, the name being obtained by naming the metal and following it by a modified name of the acid. The name of Na_2SO_4, since it is prepared from sulphuric acid, would be sodium sulphate. In the same way phosphoric acid produces phosphates, and oxalic acid, oxalates. The question then comes, why do we not call the compound made by replacing the hydrogen of hydrochloric acid with sodium, sodium hydrochlorate, instead of sodium chloride? This seeming exception to the rule arises because the name falls under two rules. We may either call the compound sodium hydrochlorate, in accordance with the rule just given, or, since it is a compound composed of two elements, sodium chloride, in accordance with our first rule. The simpler name is commonly used.

In the above cases of compounds derived from acids, you will notice that all the names of the acids end in *-ic*. Some acids have names ending in *-ous*, and compounds formed from them have names ending in *-ite*. Thus, when the hydrogen in nitrous acid, HNO_2, is replaced by sodium, we have the compound $NaNO_2$. This is called sodium *nitrite*.

Radicals explained. When the hydrogen of an acid is replaced by a metal, the other atoms do not as a rule separate, but tend to stick together and act as a single atom. Thus, if we take the hydrogen away from sulphuric acid by means of sodium, the sulphur and oxygen of the acid do not separate, but act as a single atom, forming the compound Na_2SO_4.

Such groups of atoms we call *radicals*. They are not capable of existing free and uncombined, but always exist in combination with some other element. A few of them are so important that they have been given names. One of them, OH, is called *hydroxyl*. It can combine with elements to form compounds called *hydroxides*. Thus NaOH is named sodium hydroxide. This compound has the ending *-ide*, although it has three elements in it, instead of only two, as the name would indicate.

One very important metallic radical is called ammonium,

NH_4. Like other radicals, it cannot exist alone, but its compounds are common. In many ways it acts like the metal sodium, and so we call it a metallic radical. Chemists call the compound NH_4OH ammonium hydroxide, but you know it under the name of ammonia water.

Important acid formulas. In order to name chemical compounds we must be able to recognize and name the symbols of the elements and to know the radicals of the common acids. The formulas of a few acids and salts are given below and should be memorized.

ACIDS		SODIUM SALTS DERIVED FROM THESE ACIDS	
Name	Formula	Name	Formula
		Sodium	
Nitric	HNO_3	nitrate	$NaNO_3$
Sulphuric	H_2SO_4	sulphate	Na_2SO_4
Hydrochloric	HCl	chloride	$NaCl$
Carbonic	H_2CO_3	carbonate	Na_2CO_3
Phosphoric	H_3PO_4	phosphate	Na_3PO_4
Acetic	$HC_2H_3O_2$	acetate	$NaC_2H_3O_2$

Chemical equations. It is often desirable to express the facts of a chemical change in the form of an equation. The fact that one molecule of oxygen combines with two molecules of hydrogen, thus producing two molecules of water, may be expressed in the brief equation:

$$2\ H_2 + O_2 \rightarrow 2\ H_2O$$

Chemical equations are much like those of algebra. You must have as many atoms of each element on one side of the equation as on the other. You may multiply both sides by the same quantity. You may add the same quantity to each side. You will notice, though, that we do not use an equality sign between the two members, but an arrow. This is because we wish to indicate in which direction the reaction occurs. Thus, the above equation does not mean that water becomes hydrogen and oxygen, but that hydrogen and oxygen become water.

In some cases the chemical change may take place in either direction, in which case we use a double arrow, as in the following example. Calcium carbonate (marble), when heated, will give off carbon dioxide, forming calcium oxide (quicklime). If the carbon dioxide is not allowed to escape, the calcium oxide will absorb it, and be converted into calcium carbonate. The double arrow in the equation shows that the reaction goes on in both directions.

$$CaCO_3 \leftrightarrows CaO + CO_2$$

You must not imagine that every chemical change that you can represent by an equation will take place. You must know that the chemical change involved is possible before you attempt to write the equation.

In writing equations it is necessary to balance them, that is, to see that there are the same number of atoms of each element on each side of the equation. Suppose you have prepared hydrogen by the action of hydrochloric acid upon zinc, and wish to write the equation representing the action between them. You start by putting down all the compounds used, and the compounds formed, thus:

$$Zn + HCl \rightarrow ZnCl_2 + H_2$$

Evidently this is not complete, as we have two atoms of chlorine on the right-hand side, and but one on the left. You must use two molecules of hydrochloric acid to obtain these two atoms of chlorine, and will then have two atoms of hydrogen (one molecule) left. The complete equation is:

$$Zn + 2\ HCl \rightarrow ZnCl_2 + H_2$$

It will be good practice for you to balance the incomplete equations given below:

$$NaOH + H_2SO_4 \rightarrow Na_2SO_4 +$$
$$NaOH + H_2SO_4 \rightarrow NaHSO_4 +$$
$$CaCO_3 + HCl \rightarrow CaCl_2 + CO_2 +$$
$$Al_2O_3 + H_2SO_4 \rightarrow Al_2(SO_4)_3 +$$
$$CaO + HCl \rightarrow CaCl_2 +$$

Valence explained. One of the puzzling things to beginners in chemistry is to know how many atoms of each element to include when writing the formula of a compound. For example, you have learned that copper oxide is a compound of copper with oxygen. This, of itself, is not enough to enable you to write the formula, as it might be CuO, CuO_2, Cu_2O, etc. You must know how many atoms of oxygen and how many of copper are present. A knowledge of what chemists call *valence* will often enable you to write the correct formula.

The formulas of hydrochloric acid, HCl; water, H_2O; ammonia, H_3N; and marsh gas, H_4C,[1] have been determined by actual analysis in the laboratory. An examination of these formulas shows that chlorine, oxygen, nitrogen, and carbon differ in their ability to combine with hydrogen atoms and form compounds with it. An atom of chlorine is able to hold, or combine with, only one atom of hydrogen. An atom of oxygen can hold two, an atom of nitrogen three,

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[1] The formula of marsh gas is generally written CH_4. In this discussion it is written H_4C to make it agree with HCl and H_2O. The same thing is true of ammonia. It is commonly written NH_3.

an atom of carbon four, hydrogen atoms. This power of the atom of an element to combine with, or replace, one, two, or more hydrogen atoms is called its *valence.*

Let us call the valence of hydrogen one, taking it as the standard. Chlorine will also have a valence of one, as it can combine with but one hydrogen atom. Oxygen must have a valence of two, as it will combine with two hydrogen atoms. The valence of nitrogen must be three, and the valence of carbon, four, for similar reasons.

A simple way of thinking of valence is to think of wooden blocks having one, two, or more hooks placed in them. Imagine that you have two kinds of wooden blocks, one red and the other blue (Fig. 41), each block having one hook. Should you attempt to hook red blocks and blue blocks together, it would be easy to hook one red block and one blue block together. There you would have to stop. You could not hook a second red block upon the blue one, for there would be no hook vacant on which to hang it. If, though, you had a third kind of block, a yellow one having two hooks, you could hang a red block on to each hook of the yellow block, thus forming a set of three blocks. If now you will call the blue blocks chlorine atoms, the red ones hydrogen atoms, and the yellow ones oxygen atoms, you will see how the molecules of hydrochloric acid and of water are put together.

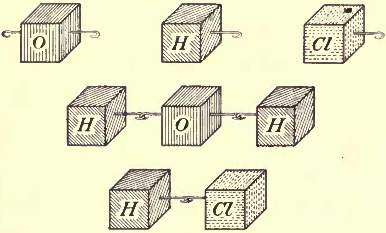


Fig. 41. — Diagram to illustrate valence.

It seems to be true that every hook of an element must be used. Carbon has four hooks. We cannot have such a compound as H_3C, as that would leave a hook unused.

You must be sure to understand that these hooks are only imaginary. We do not think for an instant that real atoms have real hooks coming from them. To the beginner, though, the question of valence is puzzling, and to think of actual blocks and hooks helps to make the matter clear. Chemists speak not of hooks but of *bonds*. In the future, then, we must not say that carbon has four hooks, but that carbon has four bonds, or, that its valence is four. We may also say that it is *tetravalent*. An element with three bonds is *trivalent;* one with two bonds, *bivalent;* and with one bond, *univalent.*

Variable valence. The subject of valence is not quite so simple as you might think from the above. Elements sometimes change their valence. Thus the valence of iron may be two or three, and we have in consequence two iron chlorides, $FeCl_2$ and $FeCl_3$. The valences of elements are given in the table on page 378.

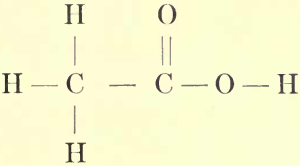
Writing formulas of salts. Having learned the formulas of some of the common acids, and the valence of some of the common metals, you can write the formulas of their compounds without trouble. Let us suppose you wish to write the formula of sodium sulphate. You know that it must contain sodium atoms, combined with the SO_4 radical. The question now is, how much of each must be taken. The formula of sulphuric acid is H_2SO_4. Since the SO_4 radical is able to hold two atoms of hydrogen, its valence must be two. The valence of sodium being one, the SO_4 radical must be able to combine with two atoms of sodium instead of two atoms of hydrogen. The formula, then, must be Na_2SO_4.

How to find the valence of an element. You can of course work this same process backward in order to find out the valence of an element, provided you know the formula of some one of its compounds. The formula of potassium phosphate is K_3PO_4. Knowing that the valence of PO_4 is three, and that three atoms of potassium will combine with it, you know that the total valence of the three atoms of potassium must be three, or that the valence of one atom must be one. Potassium must then have a valence of one or is univalent.

The formula of aluminium oxide is Al_2O_3. Since the valence of oxygen is two, three atoms of it must have six bonds. The two atoms of aluminium must therefore have six bonds, or each atom must have three bonds, that is, the valence of aluminium is three.

Empirical and rational formulas. Compounds containing carbon often have a large number of atoms in them. The formula of cane sugar (sucrose), for example, is $C_{12}H_{22}O_{11}$. Such a formula as this, giving only the number of atoms in a molecule, is called an *empirical formula.* Such a formula does not give us all of the information that we need; for if you again think of your wooden blocks with hooks, you will see that it would be possible to link such a large number of blocks together in many ways, and still have all of the hooks used.

It is not until a chemist knows, not only the number of atoms in the molecule, but how these atoms are arranged, that he really knows much about the compound. When we have found out the actual arrangement of the atoms, we can then draw a diagram that will show not only the number, but the arrangement of the atoms. Such a diagram (formula) we call a *graphic formula.* The empirical formula of acetic acid is $C_2H_4O_2$. Since acetic acid has four hydrogen atoms in it, you would expect its radical to be C_2O_2, that this radical would have a valence of four, and that the formula of sodium acetate would be $Na_4C_2O_2$. Instead of this, the acid radical is $C_2H_3O_2$, its valence one, and the formula of sodium acetate $NaC_2H_3O_2$. An examination of the graphic formula will explain why this is so.



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

Graphic Formula of Acetic Acid.

You will notice that three of the four atoms of hydrogen are united to a carbon atom, while the fourth is united to an oxygen atom. The hydrogen atom that is united to the oxygen is easily replaced, while the others are not.

We find the same thing true in the case of other organic acids. They all contain many hydrogen atoms, some combined with carbon, and some united to oxygen. We find that only those united to the oxygen, or rather those that form part of the group C–O–O–H (called the carboxyl group) are replaceable by metals to form salts. An ordinary empirical formula would not show this, hence the importance of the graphic formula. To show this difference in the hydrogen atoms, the formula of acetic acid is better written $HC_2H_3O_2$, instead of $C_2H_4O_2$.

How graphic formulas are found. To determine these graphic formulas is often a matter of great difficulty. It can often be done by breaking up a complex compound into simpler ones, studying the arrangement of the atoms in these, and then studying the way in which these simple compounds unite to form the original substance.

In many cases how the atoms are arranged is not known, and in these cases it is almost impossible to reproduce the compound in the laboratory. When the graphic formula is known, the chemist, in many cases, can go into the laboratory and make the compound, because he knows how it is built.

The manufacture of synthetic indigo was made possible only after chemists had determined the graphic formula of the natural indigo. After it could be made in the laboratory, it was years before it could be made cheaply. To-day most of the blue cloth of commerce is dyed with synthetic indigo, made not by nature in the indigo plant, but by man. It is not an imitation, but real indigo.

A similar thing is being done in the case of rubber, but the work is not yet so far advanced. It is possible now to make rubber in the laboratory, but the artificial product costs much more than the natural. It is probable, though, that in time chemists will be able to cheapen the process, and that synthetic rubber will become an article of commerce.

These synthetic products must not be confused with imitations. To determine that the flavor of vanilla is due to a substance called vanillin, to make that synthetically, and to sell it for flavoring purposes, is legitimate. But to sell vanillin, colored with caramel, and call it " genuine vanilla extract " is a fraud.

Chemical problems. A knowledge of atomic and molecular weights is indispensable to the manufacturing chemist. A simple problem taken from your laboratory work will illustrate this. You have prepared hydrogen by the action of hydrochloric acid on zinc. Suppose you wish to know how much acid will be required to dissolve 40 grams of zinc. By first writing the equation and then calculating the molecular weights, it is possible to state a proportion that will give you the desired information.

$$\begin{array}{ccc} Zn + 2\ HCl & \rightarrow & ZnCl_2 + H_2 \\ 65 + 73 & \rightarrow & 136 + 2 \\ & 138 = 138 & \end{array}$$

Factors			Products	
Weight of 1 atom of Zn		65	Zn = 65	
	H = 1		2 Cl = 71	
	C = 35.5		136	
Molecular weight of HCl	36.5		Molecular weight of $ZnCl_2$	136
Weight of 2 molecules of HCl		73	Weight of 1 molecule of H_2	2
Total weight used		138	Total weight obtained	138

The equation tells us that 65 grams of zinc will require 73 grams of hydrochloric acid to dissolve it: then 40 grams of zinc will require x grams of hydrochloric acid.

$$65 : 73 :: 40 : x$$

Solving the proportion, $x = 44.9$ grams of acid required.

SUMMARY

Atomic weight is the weight of one atom of an element compared to the weight of one atom of hydrogen.

Molecular weight is the weight of one molecule of a substance compared to the weight of one atom of hydrogen.

Valence denotes the number of hydrogen atoms that one atom of an element can combine with, or replace.

A radical is a group of atoms that, in reactions, usually acts like one atom. A radical cannot exist free.

Naming binary compounds: Compounds containing only two elements have the more metallic element named first, followed by the name of the other element so modified as to end in *-ide*.

Naming salts from -ic acids: If the name of an acid ends in *-ic*, the name of the compound made from it ends in *-ate*.

Naming salts from -ous acids: If the name of an acid ends in *-ous*, the name of the compound made from it ends in *-ite*.

Exercises

1. Find the molecular weight of $NaOH$, K_2O, $NaHCO_3$, HCl.

2. Name $NaOH$, KOH, $NaCl$, CuO, HgO, CuS, PH_3, NH_3.

3. Name CO, CO_2, P_2O_3, P_2O_5, $FeCl_2$, $FeCl_3$.

4. The valence of iron, Fe, is 3. The valence of oxygen, O, is 2. What is the formula of iron oxide?

5. The formula of potassium chloride is KCl. The formula of tartaric acid is $H_2C_4H_4O_6$. What is the formula of potassium tartrate?

6. The formula of silver nitrate is $AgNO_3$. The formula of nitric acid is HNO_3. What is the valence of silver?

7. Balance the following equations:

$$HCl + NaOH \rightarrow NaCl + ?$$
$$H_2SO_4 + Na_2O \rightarrow NaHSO_4 + ?$$
$$HNO_3 + K_2CO_3 \rightarrow KNO_3 + CO_2 + ?$$

8. How many grams of hydrogen can be obtained by decomposing 9 grams of water? How many grams of oxygen?

9. How many grams of chlorine and hydrogen will be required to combine and form 80 grams of hydrochloric acid?

10. What is the percentage composition of water? (This is the same thing as asking how many grams of oxygen and hydrogen will be required to form 100 grams of water.)

CHAPTER XI

CHLORINE

How electrolysis is used by chemists. Nature has revealed to us many of her secrets, but there are many still to be discovered. One of these is the cause of the changes that take place in the appearance and properties of elements when they unite to form compounds. A notable example of this is the compound that we all know so well, common salt. If table salt is melted in a suitable vessel, and an electric current then passed through it, the salt is decomposed into two elements, sodium and chlorine.

Neither of the two elements, sodium and chlorine, obtained from salt by electrolysis, resembles salt in the slightest. Sodium is a silvery white, soft, light metal, while chlorine is a bad-smelling, poisonous gas. You will remember that water can be decomposed in this same way. The electric current is of great use to chemists in pulling many other compounds to pieces. Many metals, as aluminium and calcium, are obtained by passing an electric current through one of their fused compounds.

The preparation of chlorine. There are easier ways of obtaining chlorine in the laboratory than by decomposing salt. One of the common acids that you have used is called hydrochloric acid, or, to use its old name, muriatic acid. This acid has the formula HCl, that is, it is composed of chlorine and hydrogen. You know that oxygen has a great liking for hydrogen. If, then, we add to hydrochloric acid some substance that will easily give up oxygen (an oxidizing agent), the oxygen will unite with the hydrogen, leaving the chlorine free. Any good oxidizing agent can be used. Potassium chlorate, hydrogen peroxide, or potassium permanganate will serve the purpose, as you will find in the laboratory. We usually use manganese dioxide, which is inexpensive. It is also insoluble, and so the action is slow, which is an advantage.

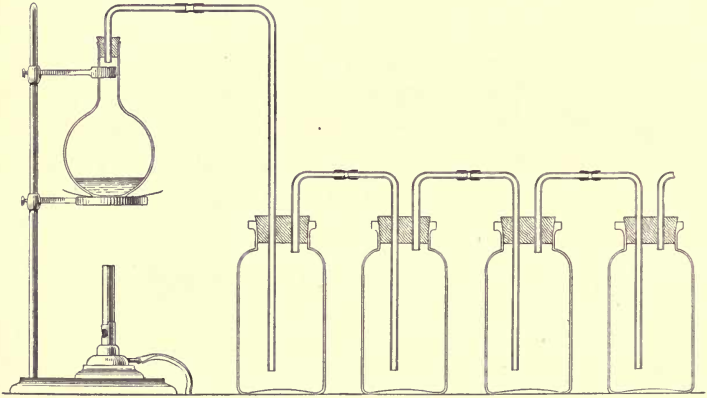


FIG. 42. — Preparation of chlorine.

To make chlorine, place manganese dioxide in a flask, and add hydrochloric acid (Fig. 42). To hasten the action heat the flask. A greenish yellow gas is given off, which is chlorine. If we bubble this gas through water, the water becomes colored, showing that the gas dissolves in it. Chlorine is therefore not collected like oxygen, for the chlorine would dissolve in the water. Chlorine is collected by displacement of air. Simply place the end of the delivery tube from which chlorine is flowing, close to the bottom of a bottle, and the chlorine will fill the bottle, driving out the lighter air. The color of the gas enables us to tell when the bottle is filled with it.

An illustration of completing a chemical equation. The formula of manganese dioxide is MnO_2. That is, one molecule contains one atom of the metal manganese and two atoms of oxygen. Only one of the oxygen atoms is easily given up, that is, only one of them can be used for oxidizing purposes. Let us write the formula, then, as $MnO(O)$, so as to show the difference between the oxygen atoms. Now we are ready to work out the equation that shows us just how much MnO_2 and HCl we must use. The one atom of oxidizing oxygen can combine with two atoms of hydrogen, and to furnish these two hydrogen atoms we must use two molecules of hydrochloric acid, or:

$$MnO(O) + 2\ HCl \rightarrow MnO + H_2O + 2\ Cl$$

The two atoms of chlorine will at once combine to form one molecule, so it would be better to write Cl_2 rather than $2\ Cl$. The MnO that is left is soluble in hydrochloric acid, and as we have plenty of this present, a second equation takes place:

$$MnO + 2\ HCl \rightarrow MnCl_2 + H_2O$$

These two changes take place simultaneously, so that we would ordinarily write only the complete equation obtained by adding these two:

$$\begin{array}{c} MnO(O) + 2\ HCl \rightarrow MnO + Cl_2 + H_2O \\ MnO + 2\ HCl \rightarrow MnCl_2 + H_2O \\ \hline MnO_2 + MnO + 4\ HCl \rightarrow MnO + MnCl_2 + Cl_2 + 2\ H_2O \end{array}$$

As the MnO occurs on both sides of the equation, we may cancel it, which leaves as the final equation:

$$MnO_2 + 4\ HCl \rightarrow MnCl_2 + 2\ H_2O + Cl_2$$

If we had used potassium chlorate, a similar change would have taken place:

$$KClO_3 + 6\ HCl \rightarrow KCl + 3\ H_2O + 3\ Cl_2$$

Preparation and use of bleaching powder. One other method of obtaining chlorine is worth remembering. You have doubtless used at home bleaching powder, called also chloride of lime. This is a white powder, obtained by passing chlorine over lime, when the lime combines with large quantities of the chlorine.

$$CaO + Cl_2 \rightarrow CaOCl_2$$

If an acid is added to the bleaching powder, the chlorine is again set free, and this is an easy way of obtaining chlorine gas in the home.

$$CaOCl_2 + 2\ HCl \rightarrow CaCl_2 + H_2O + Cl_2$$

Properties of chlorine. Chlorine is a heavy gas of a greenish yellow color, and of an exceedingly disagreeable odor. It affects the mucous membranes, as the lining of the throat and nose, producing the sensations of a bad cold. It is poisonous, and care should be taken not to inhale it. Should you suffer from its effects at any time, inhaling the vapor of alcohol will bring some relief. It is soluble in water, one volume of water dissolving two volumes of the gas, at ordinary temperature.

Chlorine combines readily with metals. This can be easily shown by heating a copper wire and plunging it into a jar of the gas. Dense yellow fumes of copper chloride will appear, and the energy of the chemical change is so great that the copper is heated to a glowing red heat.

Gold is sometimes obtained from its ores by grinding the ore to a very fine powder, and then shaking this powder with chlorine water. The gold combines with the chlorine, forming gold chloride, $AuCl_3$. This dissolves in the water, is drawn off, and the gold obtained from it.

When a lighted candle is placed in a bottle of chlorine, the flame becomes very smoky and of a peculiar reddish hue. The chlorine disappears, and a black deposit forms on the sides of the jar. If you will recall that candle wax is a compound of hydrogen and carbon, you can explain what has happened. Hydrogen easily combines with chlorine, while carbon does not. Tests show that the gas left in the jar is an acid and that the black deposit will burn. The result of the chemical change in the bottle is that hydrochloric acid is formed and that carbon is set free. Other compounds of carbon and hydrogen behave in the same way when burned in chlorine.

Chlorine can easily be condensed to a liquid, and is often sold in this form, the liquid being contained in steel cylinders that can withstand great pressure. It combines readily with hydrogen, forming hydrochloric acid. For this reason, the water solution of chlorine cannot be kept for any length of time, as the hydrogen of the water combines with the chlorine, forming hydrochloric acid and liberating oxygen.

$$2\ H_2O + 2\ Cl_2 \rightarrow 4\ HCl + O_2$$

Uses of chlorine. Chlorine has two great uses, both of which are interesting to us. It is used as a bleaching agent and as a disinfectant. Its bleaching power is shown only when the article to be bleached is wet. This is because the chlorine combines with the hydrogen of the water, leaving oxygen. This oxygen is very energetic. This is often true of elements at the instant that they are set free, and, to express this condition, we say that the element is in the *nascent* *state.* This energetic oxygen oxidizes the color that we wish to destroy. Not all colors will be bleached, but only those that can be easily oxidized. Thus, printer's ink contains carbon, and since carbon is not easily oxidized at low temperatures printer's ink is not bleached by chlorine.

Bleached cotton has been through a process at the mills as illustrated in Figure 43. The unbleached cotton is on a roll (*A*). It passes through bleaching vats (*BB*) which contain a thin paste of bleaching powder. Following each bleaching bath the cloth passes into a weak acid (*SS*). Sodium sulphite (*C*) neutralizes any chlorine in the cloth. After thorough washing in water (*WW*), the cloth is dried and ironed (*H*) and the bleached cloth rolled on a drum (*F*).

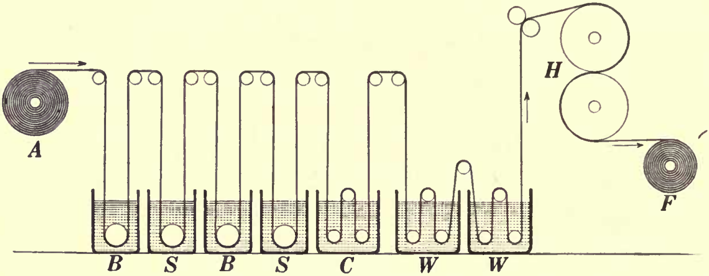


FIG. 43. — Process of bleaching cloth.

The second use of chlorine, that of a disinfectant, is dependent upon the same facts. It oxidizes the offensive matter. For this use bleaching powder is suitable. A little of the powder, scattered in a place where offensive odors are being generated, slowly decomposes and oxidizes the objectionable matter, thus destroying the odor.

Use of Javelle water in the laundry. Possibly you have never had the experience of sending a shirt or a shirtwaist to the laundry, and having it come back dazzlingly white, only to find on wearing it that the cloth has had its strength destroyed. This "tendering" of the cloth is due to the incorrect use of chlorine as a bleach. If chloride of lime is mixed with washing soda, a clear solution of sodium hypochlorite (Javelle water) results. This is often used by laundries, and in the home, as a bleach, as it gives off chlorine readily. Properly used, it is not objectionable. But if a quantity of it is poured on a tubful of clothes, some of the clothes receive more than their share, and the strength of the cloth is consequently destroyed. Some materials, as silk and wool, cannot be bleached with chlorine, as the fiber would be turned yellow and destroyed.

Ink eradicators. One of the common ink eradicators is made of sodium hypochlorite. This is rubbed on the ink spot that is to be removed, and the excess blotted off. A weak acid solution is then put on the ink. Chlorine is set free, and this bleaches the ink. Many stains will yield to the same treatment. Care must be used not to have the solution so strong that the fiber of the cloth will be weakened.

SUMMARY

Preparation of chlorine: Chlorine is prepared by the electrolysis of sodium chloride, and by the oxidation of hydrochloric acid.

Properties of chlorine: Chlorine is a heavy gas, greenish yellow in color, soluble in water, and active chemically.

Uses of chlorine: Chlorine is used for bleaching and disinfecting. Javelle water (sodium hypochlorite) is a good bleaching agent.

Exercises

1. Why does not chlorine bleach printer's ink?
2. Is Javelle water suitable to remove a stain from colored calico?
3. Will chlorine bleach all writing inks? Explain.
4. Is Javelle water a suitable bleach for a woolen sweater?

CHAPTER XII

ACIDS, BASES, AND SALTS

A study of hydrochloric acid as a typical acid. Many of the substances used in the household have a decidedly sour taste. This taste is generally due to the presence of some one of a class of bodies known as acids. Vinegar contains acetic acid, lemon juice contains citric acid, while the small green apples that occasionally cause much trouble to small boys contain malic acid. Acids differ in many of their physical properties. Hydrochloric acid, for instance, is a gas; acetic acid is a liquid; and oxalic acid is a solid. They all, however, have certain chemical properties in common. We shall now study a typical acid, hydrochloric acid, and see if we can from this study find what are the chemical characteristics of acids.

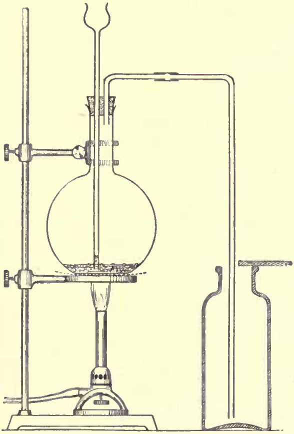


FIG. 44. — Preparation of hydrochloric acid.

Preparation. A mixture of equal volumes of hydrogen and chlorine, when exposed to sunlight, combines with explosive violence, forming a colorless gas which we name hydrochloric acid.

This is an expensive and undesirable way of forming the acid. A much better laboratory method is to place in a flask, provided with a delivery tube, a quantity of salt (Fig. 44). Then add some sulphuric acid and warm gently. A colorless gas, called hydrochloric acid, is given off in large quantities, and may be collected in bottles by displacement of air. It cannot be collected by displacement of water owing to its very great solubility in water.[1]

$$NaCl + H_2SO_4 \rightarrow NaHSO_4 + HCl \uparrow$$

The hydrochloric acid that we ordinarily use in the laboratory is a solution of this gas in water. This is a colorless liquid having a specific gravity of 1.13, and contains about 33% of real hydrochloric acid. It often goes under the name of muriatic acid.

Properties. Hydrochloric acid is corrosive in its action, dissolving the oxides of most of the metals, with the formation of chlorides, as the following equation shows:

$$ZnO + 2\,HCl \rightarrow ZnCl_2 + H_2O$$

It is too powerful an acid to be suitable for most household uses. Old iron stains, however, that resist almost every other form of treatment, may be removed from cloth by soaking the stain in a dilute solution of hydrochloric acid, and then washing the fabric well.

Hydrochloric acid, in common with all other acids, causes most dyes to change color. Should you get any of it upon your dress, a red spot may be the result. The original color usually can be restored by adding ammonia, and then rinsing with clean water.

Typical properties of acids. We have used hydrochloric acid in the laboratory to prepare hydrogen by its action on

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[1] At the room temperature 1 c.c. of water will dissolve 450 c.c. of hydrochloric acid gas weighing 0.8 gram. This gives 1.5 c.c. of concentrated hydrochloric acid solution.

zinc. We have also used sulphuric acid and magnesium, and found that they set hydrogen free. Since similar results are obtained from using some other acids and some other metals, we may conclude that the action of acids on metals is, in general, to set hydrogen free.

$$2\,HCl + Zn \rightarrow ZnCl_2 + H_2$$
$$H_2SO_4 + Mg \rightarrow MgSO_4 + H_2$$

Another property common to all acids is their action on certain colors. From one of the lichens a blue dye called *litmus* is obtained. If an acid is added to this blue litmus, the blue changes to red. If, then, we wish to know whether any liquid contains an acid, we may add a little blue litmus to it. If the litmus changes to red, we know that an acid is present. Since litmus indicates the presence of an acid, chemists call it an *indicator*.

Let us sum up the properties of acids. They are bodies having a sour taste. They all contain hydrogen, and this hydrogen is usually united to nonmetallic elements. Metals can replace this hydrogen, setting it free. They all turn blue litmus red. They are generally soluble in water.

Typical properties of bases. Another class of bodies has properties which are opposite to those of acids. We call this class *bases*. They have a bitter or alkaline taste, they turn red litmus blue, and feel soapy to the touch. Chemically they are hydroxides of metals. One of the most important bases is sodium hydroxide. A solution that turns red litmus blue is said to have an *alkaline reaction*. Strong soluble bases, like sodium hydroxide, are often called *alkalies*.

Sodium hydroxide. Soda lye, or sodium hydroxide, $NaOH$, is one of the sodium compounds familiar to housekeepers. It has the power of combining with almost all forms of grease to form soap, and so is of use in cleaning out greasy sinks and pipes, being more powerful in its action than the washing soda so commonly used for this purpose. It must, however, be handled with some care, as its action on the skin is corrosive. It is a white, deliquescent substance.

It may be prepared by the action of sodium on water.

$$2\ Na + 2\ H_2O \rightarrow 2\ NaOH + H_2$$

Sodium hydroxide is prepared on a commercial scale from quicklime which is mixed with water until a milklike fluid is obtained. To this a solution of washing soda is added. The two chemicals react, and the calcium carbonate that is formed, being insoluble, settles to the bottom of the tank. The sodium hydroxide solution is then drawn off and evaporated, leaving a white, deliquescent solid.

$$CaO + H_2O \rightarrow Ca(OH)_2$$
$$Ca(OH)_2 + Na_2CO_3 \rightarrow CaCO_3 + 2\ NaOH$$

Much sodium hydroxide is manufactured by the electrolysis of common salt and at the same time chlorine is produced. Large quantities of sodium hydroxide are used in the preparation of soap.

Neutralization and the formation of salts. When an acid and a base are mixed in solution, chemical action takes place. The hydrogen of the acid combines with the OH (hydroxyl) of the base, and water is formed. The metal of the base and the acid radical then combine and form a body known as a *salt*. Thus, sodium hydroxide, mixed with acetic acid, forms water and sodium acetate.

$$NaOH + HC_2H_3O_2 \rightarrow H_2O + NaC_2H_3O_2$$

You must carefully distinguish between *salt*, which is sodium chloride, and *a salt*, which is one of a class of bodies.

When an acid and a base combine, the characteristic properties of both are destroyed. The new substances are usually *neutral* to the litmus test, and their taste is neither acid nor alkaline. This chemical change is called *neutralization.*

Not only the hydroxides, but also the carbonates of metals will neutralize acids. If sodium carbonate, called washing soda, is mixed with an acid, carbon dioxide and water are formed, and a salt is left. If hydrochloric acid is used, sodium chloride will be formed.

$$Na_2CO_3 + 2\,HCl \rightarrow 2\,NaCl + H_2O + CO_2$$

Exact neutralization. To neutralize an acid exactly by means of a base requires some care, but it can readily be done with the aid of litmus. If to hydrochloric acid in a flask a few drops of litmus solution are added, the liquid becomes red in color. If now a solution of sodium hydroxide is added, shaking well after each addition so as to mix the liquids thoroughly, the acid and base will combine, forming water and sodium chloride. When just enough sodium hydroxide has been added to exactly neutralize the acid, the litmus will change to a violet color, showing that neither acid nor base is present in the solution. If the water is now evaporated until the solution is saturated and the dish is then set aside, the sodium chloride will crystallize.

That sodium chloride is really obtained is shown, first, by the shape of the crystals, which is cubical, the same as that of crystals of salt purchased in the grocery store. Second, they have the well-known salty taste of common salt.

Composition of baking powders. The baking powders so commonly used will also serve to illustrate neutralization. The common cream of tartar baking powder is composed of potassium hydrogen tartrate (cream of tartar), which has certain acid characteristics, and sodium bicarbonate. When these two are mixed, the acid part of the cream of tartar is neutralized and a new substance, a salt, sodium potassium tartrate (Rochelle salts), is formed. At the same time, a gas, carbon dioxide, is given off, and it is this gas that makes the cake light.

$$KHC_4H_4O_6 + NaHCO_3 \rightarrow KNaC_4H_4O_6 + H_2O + CO_2 \uparrow$$

Chemically, there is no reason why any acid should not be substituted for the cream of tartar.

A baking powder made of hydrochloric acid and sodium bicarbonate would be satisfactory, except that it would be difficult for the cook to mix the two in such proportions that she could be sure that neither would be present in excess. Hydrochloric acid and washing soda, in the same way, could be used, but would be open to the same objection.

Soda biscuits. Those of you who have eaten old-fashioned soda biscuits, leavened by the use of sour milk and sodium bicarbonate (saleratus), know that sometimes the biscuits are sour. You can now readily see why. The sourness of milk is a variable quantity, and no receipt can give the amount of bicarbonate necessary to neutralize these varying amounts of acid. Sometimes the biscuits show yellow streaks and have a somewhat soapy taste. This is due to an excess of saleratus, caused by a deficient acidity in the milk.

Replacement of one element by another. In the same way that a metal will replace the hydrogen of an acid, other metals may be used to replace the first. Thus a piece of zinc placed in a solution of lead nitrate will replace the lead, forming lead and zinc nitrate. Similarly copper will replace the silver in silver nitrate. The more metallic element will replace the less metallic. Arranging a few of the common metals in the order of their metallic properties will tell us how they will replace one another. Magnesium, aluminium, zinc, lead, copper, silver, and gold form such a series. Each metal will replace all that come after it. Those that follow are less metallic than those which precede.

SUMMARY

Hydrochloric acid is made by the action of sulphuric acid on sodium chloride. The water solution of it is in commercial use.

An acid is a substance containing hydrogen. As a rule the hydrogen is easily replaced by a metal. Acids are sour, and turn blue litmus red.

A base is the hydroxide of a metal. Bases are caustic, and turn red litmus blue.

A salt is the product of the action of an acid on a base. A salt may be regarded as an acid in which the hydrogen has been replaced by a metal, or, as a base in which the hydroxyl has been replaced by an acid radical.

Neutralization is the combining of an acid and a base, forming water and a salt.

Sodium hydroxide is made by the reaction of sodium carbonate and calcium hydroxide. It is also made by the electrolysis of common salt.

Exercises

1. Name two bases used in the home.

2. Name three acids used in the home.

3. Why will a washing compound often change the color of cloth?

4. Why is ammonia added to the water used in washing windows?

5. Why is it necessary to keep baking powder in air-tight cans?

6. Given a colorless liquid, how could you tell whether it contained an acid or a base?

CHAPTER XIII

SODIUM AND ITS COMPOUNDS

Properties of sodium. You will remember that, by decomposing fused common table salt by electrolysis, a yellowish green gas called chlorine and the metal sodium are obtained.

You doubtless think of a metal as being a heavy, hard substance. This metal sodium, however, has a character entirely different from that of an ordinary metal. It is not hard, but so soft that it can be molded in the fingers like putty; nor is it heavy, in fact its specific gravity is so small that it will float upon water. Then, too, we ordinarily think of metals as being exceedingly permanent bodies, not easily corroded or destroyed. This metal sodium, however, corrodes, or rather, oxidizes, so rapidly in the air that it is only possible to keep it in the metallic state by preventing the air from coming in contact with it.

Evidently, then, a chemist has an entirely different idea of a metal from that ordinarily held. To a chemist, sodium is a typical metal, because it combines so energetically with oxygen, reacts so strongly with acids, and because its hydroxide is so strongly alkaline.

We will sum up, then, the properties of sodium. It is a soft metal of a silvery color, very light, having a specific gravity of .935, a melting point of 207.7° F., an atomic weight of 23, is easily oxidized, and conducts heat and electricity very well. It is kept under kerosene or some other liquid that contains no oxygen.

Action of sodium on water. The action of sodium upon water is very interesting. If we fill a test tube with water, invert it in an evaporating dish of water, and then quickly thrust a piece of sodium the size of a small pea under the mouth of the test tube, we see that the sodium melts and spins around actively, while the water in the test tube disappears (Fig. 45). If, now, we place a finger over the mouth of the test tube, and bring it to a light, we find that the tube is filled with an inflammable gas which we recognize as our old acquaintance, hydrogen. If we wet our fingers in the water in the evaporating dish and rub them together, they feel soapy or slippery. On evaporating a little of this water, a white solid is obtained, very different from the sodium with which we started. Evidently the sodium in setting the hydrogen free has combined with part of the water.

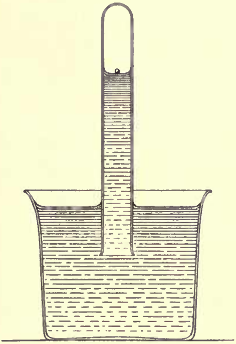

FIG. 45. — Action of sodium upon water.

$$2\,Na + 2\,H_2O \rightarrow 2\,NaOH + H_2$$

Sodium chloride or table salt. Compounds of sodium are found widely distributed in nature, many rocks and soils containing them. Several natural compounds occur almost pure, and are of considerable economic importance. The most useful of these is sodium chloride, or common table salt. This occurs in very many localities in the world in large quantities. In the United States, it is found in enormous quantities in Michigan, Utah, and New York.

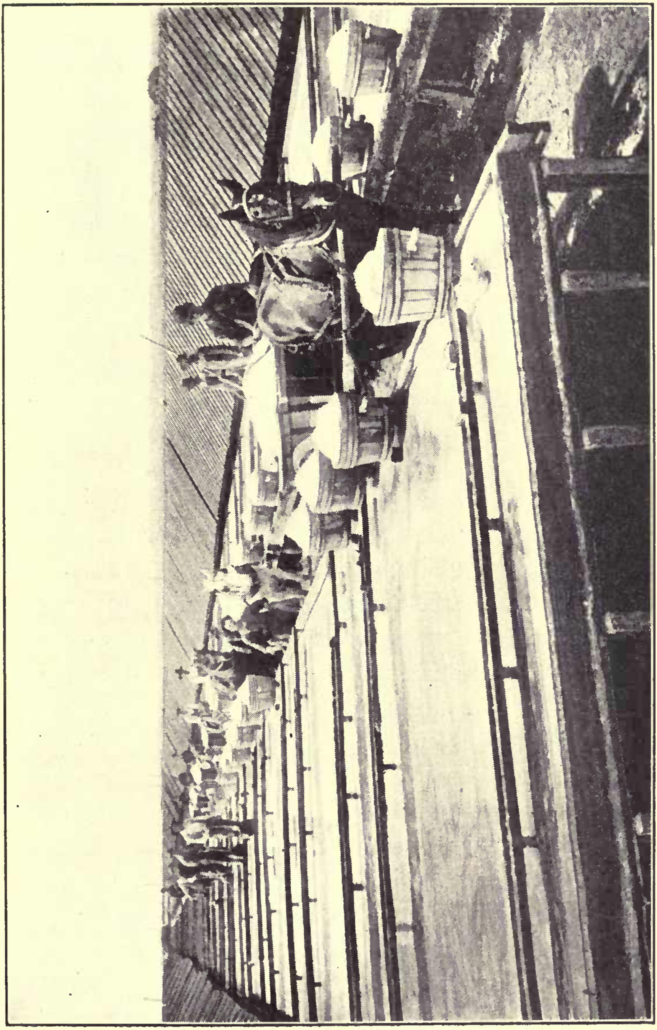

FIG. 46. — Evaporation tanks in a salt works.

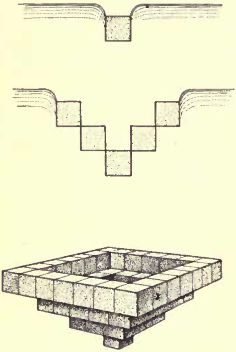

FIG. 47. — Hopper shaped crystals of salt.

Salt is commonly extracted from the ground by sinking two wells a short distance apart. Down one of these wells water is pumped. This water dissolves the salt, forming a strong brine. This brine is then pumped up the second well, and some of the water evaporated (Fig. 46). The salt crystallizes, the mother liquor is drained off, and the crystals of salt dried. Thus our common salt is obtained. The size of the salt crystals is determined by the degree of concentration of the brine and the rapidity of the crystallization. The more slowly the crystals are formed, the larger they are. It will be interesting for you to determine what the crystalline form of salt is, by examining a little under a common magnifying glass (Fig. 47).

Rock salt is salt that is actually mined from the ground, just as is coal or iron ore. Its usual reddish brown color is due to the presence, as an impurity, of a small amount of iron oxide. Rock salt sometimes occurs in perfectly transparent masses which are used by scientists to make lenses and prisms.

Rock salt has been formed by the slow evaporation of large bodies of salt water and the subsequent covering up of the deposit. This same process is now going on in our country in the Great Salt Lake. The water of this lake is a saturated solution of salt. The lake itself is gradually drying up, and the salt is being deposited upon its shores as a glistening white layer some inches in thickness. If this process continues, we shall have a bed of rock salt similar to those that exist in Siberia and Austria.

The ocean is our great reservoir of sodium chloride. It is calculated that 36,000,000,000,000,000 tons of salt exist in the ocean, and much salt is obtained from it. At Turks' Island, for example, the ocean water is run into shallow pans, and the water evaporated by the heat of the sun. The resulting large, coarse crystals are largely used in freezing ice cream.

Prepared table salt. One of the minor bothers of the housewife is the tendency of salt to absorb water, and to stick in the salt cellars. It is commonly said that salt is hygroscopic, that is, gathers water from the air. This is not true. It is the presence of magnesium chloride, $MgCl_2$, a deliquescent substance, and one of the common impurities of salt, that causes this moistening of the salt. It can, therefore, be overcome by carefully purifying the salt, but, as this is expensive, another method is commonly resorted to. A small quantity of starch, or, better, precipitated chalk, is mixed with the salt which has previously been carefully dried. This starch coats over the little particles of the salt, and prevents their sticking together.

Uses of salt. The use of salt as a seasoning in food, and as a preservative for meats, is so well known that no details need be given. It has another use in our diet. The gastric juice contains a small amount of hydrochloric acid, and the chlorine for this comes from the salt that we consume. Large quantities of salt are also used in the preparation of sodium carbonate, or washing soda.

Preparation of washing soda. Sodium carbonate, Na_2CO_3, or washing soda, is prepared by the action of ammonium hydroxide and carbon dioxide upon a solution of sodium chloride. This is known as the Solvay process.

$$NH_3 + NaCl + CO_2 + H_2O \rightarrow NaHCO_3 + NH_4Cl$$

The sodium bicarbonate, being insoluble in the liquid, is deposited as a crust on the large cylindrical tanks in which the operation is carried out. The sodium bicarbonate is filtered out, and then heated. This drives off the carbon dioxide which is used in the preparation of more of the bicarbonate.

$$2\ NaHCO_3 + \text{heat} \rightarrow Na_2CO_3 + H_2O + CO_2 \uparrow$$

The resulting mass is dissolved in water and recrystallized. Sodium carbonate is noteworthy because, when it crystallizes, each molecule combines chemically with ten molecules of water; that is, the formula of crystallized sodium carbonate is not Na_2CO_3, but $Na_2CO_3 \cdot 10\ H_2O$.

The clean, glassy crystals of washing soda, then, that you buy from the grocer, are more than half water by weight, or, to be exact, 286 pounds of this crystallized washing soda contains 180 pounds of water. On standing, these glassy crystals lose most of their water of crystallization, and are converted into a dull white powder.

When, then, you buy washing soda, do not too carefully insist upon receiving the glassy crystals, but rather choose the white powder that forms in the bottom of the barrel, as you will thus receive twice as much for your money, a thing much to be desired in this day of high prices.

The great use for this washing soda in the household is in softening hard water and as a cleaning agent. Practically all washing compounds contain it and you are doubtless familiar with its use in cleaning greasy frying pans.

Sodium bicarbonate, or baking soda. Not all of the sodium bicarbonate is converted into washing soda; much of it is sold and consumed under the name of baking soda. This substance goes under many other names, a few of which are: saleratus, cooking soda, sodium hydrogen carbonate, and sodium acid carbonate. The white powder is a mild alkali, and is used extensively in the manufacture of baking powders. It is one of the common household remedies for sour stomach.

Borax. Borax, $Na_2B_4O_7 \cdot 10\ H_2O$, sodium tetraborate, is a white crystalline substance that contains ten molecules of water of crystallization. It is used in soaps as a mild alkali. It is also used in some washing compounds.

Water glass, or sodium silicate. Sodium silicate, or water glass, Na_2SiO_3, is another sodium compound that is used in the home. It is a clear, hard solid, that dissolves slowly in water, making a thick sirup, the form in which it is usually sold. This forms an excellent cement, as, when it has once solidified, it is very strong and difficult to dissolve.

A thin solution of water glass is often used to preserve eggs. It fills up the tiny pores in the shell of the egg, and thus prevents the water in the egg from evaporating and also prevents any germs that may be in the air from gaining access to the contents of the shell. It is also used as a filler in laundry soap.

Other sodium compounds. Several other compounds of sodium are of industrial importance. Sodium nitrate, $NaNO_3$, or Chili saltpeter, is a white crystal found in Chili, and is used in large quantities as a fertilizer. Sodium thiosulphate, $Na_2S_2O_3$, the familiar hypo of the photographer, is a solvent for silver bromide, and is used in fixing photographic plates and papers.

General properties of the sodium salts. The sodium salts are practically all soluble in water, and form white crystalline substances. They are usually stable. Owing to their wide distribution, their cheapness, and their solubility, they are widely used both in the home and in the industries.

Potassium and its compounds. The metal *potassium* is very much like sodium. Whatever sodium will do chemically potassium will also do, but do it more energetically. Thus, sodium decomposes water. So does potassium, but more violently. Sodium oxidizes in the air. So does potassium, but more easily. Many potassium salts are of importance. We cannot, however, spare the time to study them. It will be enough if you remember that, generally speaking, potassium forms the same classes of salts as sodium, and that they have much the same properties. The main difference between sodium and potassium salts is that the potassium compounds are usually more soluble.

SUMMARY

Sodium is prepared by the electrolysis of fused sodium hydroxide.
Washing soda is sodium carbonate.
Baking soda is sodium hydrogen carbonate.
Table salt is sodium chloride.
Potassium is much like sodium, and forms similar salts.

Exercises

1. If sodium is light and soft, why do we call it a metal?
2. How many names can you give to $NaHCO_3$?
3. Of what is baking powder composed?
4. Name three sodium compounds that you use in large amounts.
5. Name one potassium compound that is used in the home.
6. Could you obtain salt from sea water at home?

CHAPTER XIV

AMMONIA AND AMMONIUM COMPOUNDS

Preparation of ammonia. Whenever organic material containing nitrogen compounds decays, a gas called ammonia, NH_3, is set free. This is not a suitable method to use for obtaining it in the laboratory. A better method is to place in a test tube a mixture of ammonium chloride and some base, as sodium hydroxide (Fig. 48). On warming, the gas ammonia is given off in large quantities, and can be collected by displacement of air. Any ammonium compound and any base can be used.

$$NH_4Cl + NaOH \rightarrow NH_3\uparrow + H_2O + NaCl$$

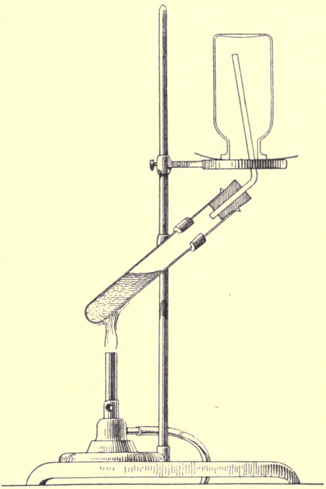

FIG. 48.—Preparation of ammonia.

Properties of ammonia. The gas thus obtained is colorless, and has a very pungent characteristic odor. Since ammonia contains hydrogen, you might expect it to burn. If we try the experiment of directing a stream of the gas coming from the delivery tube against a Bunsen burner flame, you will notice that the ammonia gas burns as long as it is in the

Bunsen burner flame, but that when removed, it goes out. That is, ammonia will burn only so long as we supply it with heat. It will not burn under ordinary conditions.

It is very soluble in water. At 50° F., one quart of water will dissolve 670 quarts of the gas. This solution is not merely physical, but a new compound is formed, ammonium hydroxide.

$$NH_3 + H_2O \rightarrow NH_4OH$$

The colorless solution has a sharp, burning taste, and smells of ammonia. It has many names: ammonium hydroxide, ammonia water, aqua ammonia, and spirits of hartshorn. It is an unstable compound, heat easily breaking it up into ammonia and water. This suggests a quick method of obtaining ammonia in the laboratory, when we wish a little of the gas.

$$NH_4OH + \text{heat} \rightarrow NH_3 \uparrow + H_2O$$

The "household ammonia" that you buy for cleaning purposes is a weak, impure, ammonium hydroxide, obtained as a waste product in certain manufacturing operations. You will get much more for your money if you will buy a bottle of the concentrated aqua ammonia, and dilute it as you need it.

Commercial production of ammonia. Commercially, ammonia is made by the destructive distillation of soft coal. Since soft coal is a product of the partial decay of wood in the earth, it contains hydrogen and nitrogen. When it is heated out of contact with the air, illuminating gas, coal tar, and ammonia are formed. The ammonia is absorbed in water, and from this "gas liquor" the ammonia of trade is obtained. Ammonia gas is cooled and compressed until it changes to a liquid, in which form it is sold in iron cylinders. It is used largely in the manufacture of artificial ice (p. 93).

The radical ammonium. If you compare the formulas of sodium and ammonium hydroxides, $NaOH$ and NH_4OH, you will notice that the group of atoms NH_4 takes the place of the sodium atom. This NH_4 group we call a *radical*, because in reactions it tends to stick together and act like one atom. Since it takes the place of a metal, we call it a *metallic radical*. Since it forms the compound NH_4OH, its valence must be one. Notice that the ending *-ium* is given to metals only, as sodium, potassium, and aluminium. Since NH_4 plays the part of a metal, it is called ammon*ium*. Be careful not to confuse the gas ammonia with the radical ammonium. Radicals never occur free, but exist only in compounds.

Ammonium salts. Ammonium forms many salts, as ammonium sulphate, ammonium chloride, and ammonium nitrate. Two ammonium salts are much used in the home.

Ammonium chloride, formed by neutralizing ammonium hydroxide with hydrochloric acid, is a white, crystalline body, often called sal ammoniac.

$$NH_4OH + HCl \rightarrow NH_4Cl + H_2O$$

It is the material used in the common wet battery used to ring bells. The next time your door bell does not ring, go to the cellar and examine the battery. Perhaps the water has evaporated, in which case fill it up, and in a few hours it will be in order again. If this is not the cause of the trouble, disconnect the battery, and wash out the old solution. Buy five cents' worth of sal ammoniac, dissolve it in a little water, and place the solution in the jar. Fill it up with water, again connect the wires, and probably the bell will ring again as before. Ammonium chloride is one of the substances used in the dry cell, which is now largely replacing the wet battery. Ammonium chloride is also used in soldering.

Ammonium carbonate, or sal volatile, is prepared by heating a mixture of calcium carbonate and ammonium chloride. It is a white, fibrous mass, that slowly decomposes in the air, giving off ammonia. It is this property that makes it useful to you. A lump placed in a bottle, with the addition of a little lavender water, forms the smelling salts that are so refreshing.

Ammonium salts resemble those of sodium. They are easily identified because they all give off ammonia when treated with a base, and they all sublime when heated.

SUMMARY

Preparation of ammonia. Ammonia is prepared by the action of a base on an ammonium salt, or by heating ammonium hydroxide.

Properties of ammonia. Ammonia is a colorless gas, very soluble in water forming a base, burns only when supplied with additional heat, and has a characteristic odor.

Ammonium is a metallic radical and forms salts like those of sodium.

Ammonium chloride is used in bell batteries and in soldering.

Ammonium carbonate is used in smelling salts.

Exercises

1. Is it economical to buy "household ammonia"?

2. Can you prepare smelling salts at home?

3. Name NH_4NO_3, $(NH)_2 SO_4$.

4. Some soaps are said to contain ammonia. From what you know of ammonia, do you think they can contain enough to be of any use? Explain.

5. Why do decaying leaves smell of ammonia?

6. How can you tell which of two samples of ammonia water is the stronger?

CHAPTER XV

METALS

Metals and nonmetals distinguished. When we attempt to define the word metal, we encounter certain difficulties. You would perhaps say that a metal is a hard, heavy body, having a metallic luster. But sodium, which chemically is an exceedingly good metal, is soft, lighter than water, and, as usually seen, has no luster. Iron pyrites, or fool's gold, one of the sulphides of iron, is hard, heavy, and has a strongly marked metallic luster, but it is not a metal. You might add to your list of metallic properties, that metals conduct heat and electricity well. This is true, but some nonmetals conduct heat and electricity as well as some metals. These physical properties are not a satisfactory basis for an exact definition of the word metal.

Chemically, metals are easier to define. When oxygen was studied, you learned that sodium oxide, when dissolved in water, formed sodium hydroxide, and gave a solution that turned red litmus paper blue. Sodium hydroxide is a base. Magnesium, potassium, calcium, and other metals act in the same way, that is, their hydroxides are bases. We also know that all of these metals form salts. We may, then, from a chemical standpoint, easily give a satisfactory definition. A metal is a substance whose hydroxide is a base, and which forms the positive part of a salt. We may add to this certain physical properties. Metals are solids, with the exception of mercury. They can all be obtained in a crystalline form, conduct heat and electricity well, and are generally malleable.

It is not always possible to draw a sharp line between metals and nonmetals. Some elements, as arsenic, have properties intermediate in their nature between those of metals and nonmetals. If we are thinking of arsenic as compared to chlorine, it is a metal. But, compared to sodium, it is a nonmetal. Thus we have arsenic chloride, $AsCl_3$, but sodium arsenate, Na_3AsO_4.

The general methods of obtaining metals. With the exception of a few metals, as copper, gold, and silver, metals do not occur in nature in the metallic form. The most important of their naturally occurring compounds are the oxides, carbonates, and silicates. Metallurgy is the art of extracting the metal from these compounds or ores. There are two main methods: reducing the metal from its compounds by the use of carbon, and by electrolysis. As the electric current has been available only within recent years, you may be sure that any metal known to the ancients, as iron, is either found free in nature, or is obtained by reduction with carbon; while such metals as aluminium and sodium, that have only recently come into use, cannot be obtained by reduction with carbon, but are obtained by electrolysis or by some difficult chemical process.

Metallurgy of iron. The metallurgy of iron illustrates the method of obtaining metals by reduction with carbon. Iron compounds occur everywhere in soil. The red color of ordinary brick is due to the presence of small amounts of iron oxide, and common red roofing paint is largely natural iron oxide. In the Lake Superior region of the United States, large deposits of iron oxide occur as the mineral hematite,

Fe_2O_3. This ore is mined and carried to a blast furnace, where it is heated with carbon, and so reduced to metallic iron. The carbon necessary is usually used in the form of coke, produced by heating soft coal without access of air. If the iron ore contains nothing but iron oxide, and the coke nothing but carbon, all that is necessary is to heat the two together. The coke combines with the oxygen of the iron oxide, and pure iron is obtained.

$$2\,Fe_2O_3 + 3\,C \rightarrow 4\,Fe + 3\,CO_2$$

However, both ore and coke contain certain impurities, and so it is necessary to add something that will combine with these impurities and produce a fusible mass called a slag. This third addition is called a flux. If these impurities consist of metallic oxides or carbonates, silica is used as a flux. In the heat of the furnace, all these combine to form a silicate, a fusible glass, and this, being lighter than the metal, floats upon the molten iron in the lower part of the furnace.

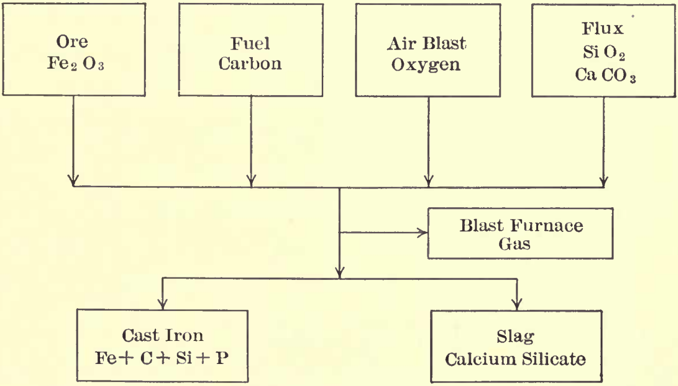

FIG. 49.— Diagram showing raw materials and products in manufacture of cast iron.

The blast furnace. To obtain iron from the ore, we must have a blast furnace in which to heat it. This consists of a huge brickwork, one hundred or more feet in height, and twenty feet through at its widest part. In shape, it is something like two cones put together base to base. Near the bottom there are two small openings at different levels, through which the iron and the slag may be drawn off. While the furnace is in operation, these are kept closed by plugs of fire clay. There are also openings through which air is blown into the furnace. The top of the furnace is closed in such a way that materials may be introduced through what is really a trapdoor.

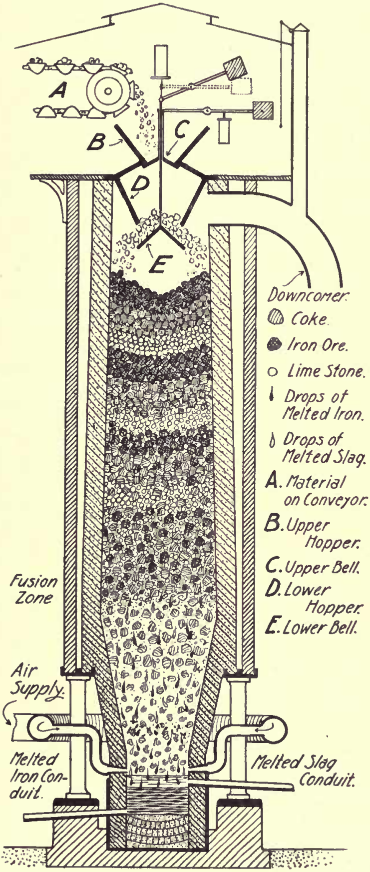

Fig. 50. — A blast furnace.

The ore, fuel, and flux are put in at the top, and sink slowly through the furnace. At the bottom, hot air is blown in. As the coke burns, it forms carbon monoxide, and this acts upon the iron oxide, reducing it, and forming metallic iron. At the same time the impurities and flux combine to form the fusible glass or slag. The molten iron and the slag drop to the bottom of the furnace, where they separate, the slag floating on the top of the iron. When a sufficient quantity has accumulated, the plugs of fire clay closing the openings are knocked out, and then the slag and iron flow out. The iron is made to flow into shallow troughs, made in a bed of sand. Here it solidifies in the form of bars called "pigs," hence the name, pig iron. The slag is sometimes used in making cement, or as ballast for a railroad track.

The pig iron or cast iron produced in this way is very impure, containing perhaps 3% of carbon and 3% of silicon, as well as small amounts of sulphur, phosphorus, and manganese. It gets its name of cast iron from an extensive use that is made of it. Since it is both cheap and fusible, it is used extensively in the form of castings for iron fences, radiators, posts, and numerous other appliances in everyday use. It is brittle, and so cannot be used where a bending strain or a sudden blow has to be resisted. It is, however, strong in resisting compression. It is made in very large amounts, as will be seen from the table below.

PRODUCTION OF PIG IRON OF WORLD FOR 1912

	LONG TONS		LONG TONS
United States	29,726,937	France	4,870,913
Germany	17,586,521	Other Countries . .	11,542,599
Great Britain	8,839,124	Total	72,566,094

Making wrought iron. Cast iron is the starting point for the manufacture of both wrought iron and steel. To make wrought iron, the cast iron is placed in a reverberatory furnace. This consists of a horizontal bed, with a low roof curved so that the heat and flames are deflected down upon whatever material is placed on the bed. The bed of the reverberatory, or puddling furnace as it is called, is first lined with iron oxide, the cast iron is then placed on it, the fire started, and the melting begins. The melted iron is stirred, and carbon, silicon, and other impurities are slowly burned out. Almost pure iron results. Pure iron is much less fusible than the impure cast iron, so that, as the end of the operation approaches, the mass becomes pasty. This pasty mass is then squeezed between rolls so as to press out the slag.

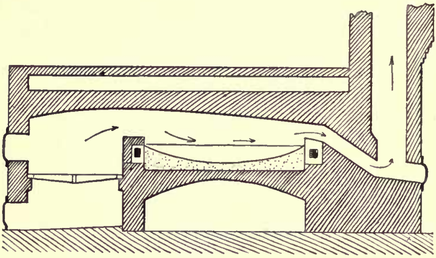
Fig. 51. — Reverberatory furnace.

To make the wrought iron uniform in composition, it is rolled into bars, these are piled one on another, and, after heating, again rolled into bars. The wrought iron thus produced is a tough, fibrous material, which can be welded.[1] It is used where it is necessary to make iron parts by forging.[2] Its freedom from carbon is believed to be the reason why wrought iron resists corrosion so successfully. It would be used more extensively, were it not for its comparatively high cost. Its cost is high because it is made in small batches of perhaps 600 pounds, hand labor is required, and the time required to make each batch is considerable.

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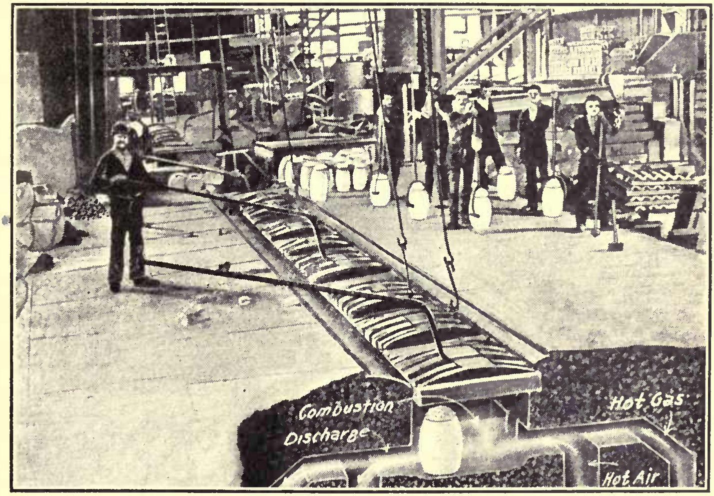
[1] If two pieces of wrought iron are heated red hot, placed one over the other, and then hammered, they stick together, and form one piece. The operation is called *welding*. It is in this way that the blacksmith makes a wagon tire.

[2] If a piece of wrought iron is heated red hot, it becomes soft, and can be hammered into any shape that the blacksmith wishes to make. This is called *forging*.

Making crucible steel. From this wrought iron the best grade of steel is made. Small pieces of wrought iron, about 90 pounds, are placed in a crucible together with carbon in the form of charcoal. When this crucible is heated in a furnace, the iron absorbs the carbon, melts, and is changed into crucible steel (Fig. 52). This is the material from which high-grade knives, springs, and things that must be both hard and springy are made. It is made in small quantities but it is the best grade of steel.



Courtesy of Crucible Steel Company of America.

Fig. 52. — Crucible steel furnace.

Bessemer and open-hearth steel. An easier way to burn out the impurities from cast iron is to blow air through molten pig iron. This product is called Bessemer steel. Ten to twenty tons of molten cast iron are placed in a large vessel, called a converter, lined with an infusible fire clay. The bottom of this converter has many holes, each of which is about the size of a lead pencil (Fig. 53). Through these holes air is blown, and in about twelve minutes the impurities are burned out. After the carbon has been burned out, an alloy of iron containing a definite amount of carbon and manganese is added. This produces a steel with a known amount of carbon.

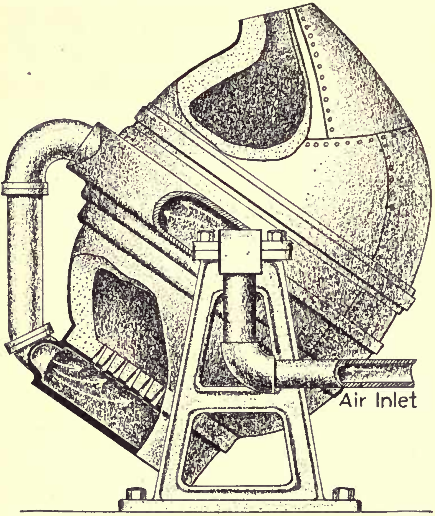


Fig. 53. — Bessemer converter.

In still another method, called the open-hearth process, pig iron is melted, together with scrap iron, in a large, low, square room, heated by burning gas (Fig. 54). The furnace hearth is lined with compounds of silica or with dolomite, depending upon the kind of iron used. Here the impurities are slowly burned out. When the carbon has been reduced sufficiently, the process is stopped and open-hearth steel results.

How steel is hardened. When steel is heated and then suddenly cooled by plunging it into water, it becomes very hard. This is the process used to harden needles and knife blades.

This hardening process leaves the steel so brittle, that it breaks like glass. If, however, it is slightly reheated, it becomes softer and less brittle. By regulating the temperature to which we reheat the steel, we can give it any desired hardness. The higher the temperature to which it is heated, the softer the steel becomes. This is called "drawing the temper."

Iron and steel contrasted. Wrought iron is the purest form of iron, and cast or pig iron the most impure. Steel is intermediate in composition.

The iron and steel industry is a good example of the large size of some of our modern industries. Furnaces in many cases have a production of six hundred tons of pig iron a day. To produce a ton of pig iron requires about two tons of ore, one half a ton of flux, one ton of carbon, and four tons of blast. About one half ton of slag and six tons of waste gases are formed.

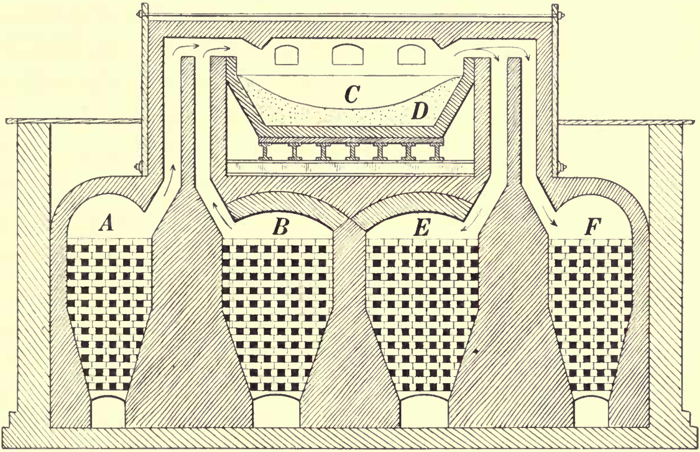


FIG. 54. — Open-hearth furnace. *A*, *B*, Air and gas heated by passing over hot bricks. *C*, Steel in process of making. *D*, Hearth lining. *E*, *F*, Bricks heated by gaseous products. These are later made inlets to warm the entering air and gas.

How iron is protected from rust. Pure iron is a soft, ductile, white metal. It readily dissolves in acids. The purest form of iron that you are familiar with is probably the thin wire that florists use in wiring flowers. In damp air, iron rusts easily, and for that reason iron kitchen utensils are often protected by coating the iron with some metal, as nickel, zinc, or tin, that does not readily oxidize. Iron buildings and bridges are protected from corrosion by painting them.

Galvanized iron. Galvanized iron is iron coated with zinc to prevent it from rusting. To make a galvanized-iron pail, a sheet-iron pail is dipped into acid to clean it, and is then washed and dried. It is then dipped into melted zinc. This forms a coating over the iron, and, since zinc does not corrode in the air, the pail is made much more durable.

In using galvanized articles, care should be taken not to bend them, as this may crack the coating of zinc. The iron is then exposed to the air, rusts, and soon the pail has a hole in it.

Tin ware. Tin is another of the metals used extensively in the home. The tin pans that you use are really not tin, however, but sheet iron coated with tin. The sheet iron is first carefully cleaned, and then dipped into a bath of molten tin. The plates are then passed between steel rollers, set close together, so as to squeeze off as much of the tin as possible. This leaves a thin coat of tin covering the iron. As tin does not corrode in the air, it serves to protect the iron underneath, while the iron gives strength and stiffness to the utensil.

Tin and lead alloy readily, as is seen in solder. The tin foil that is usually used is really made from an alloy of tin and lead.

Nickel plate. Other metals are used to cover iron and protect it. Nickel is often used, as in the handles of stove doors. It is plated upon the iron by using a bath of nickel ammonium sulphate, making the object to be plated the cathode,[1] and using a plate of nickel for the anode. On passing an electric current through the solution, the nickel is taken from the anode and deposited on the iron as a thin protective coat.

Manufacture of aluminium. The production of aluminium is typical of the methods used for obtaining metals by electrolysis. The starting point is the mineral bauxite, which is the hydroxide of aluminium. This, when heated, is changed into the oxide. As the electric current will not flow through the solid oxide, it is necessary to dissolve it. Aluminium oxide is insoluble in water, but readily dissolves in the mineral cryolite.[2]

A box lined with carbon serves as the cathode, and large carbon rods as the anode (Fig. 55). In this box the fused cryolite, containing the aluminium oxide, is placed. The current in passing through the solution decomposes it, and aluminium is set free at the cathode. Oxygen is set free at the anode, combines with the carbon, and escapes as carbon dioxide. The temperature of the bath is so high that the aluminium melts and drops to the bottom of the box, whence it is drawn from time to time.

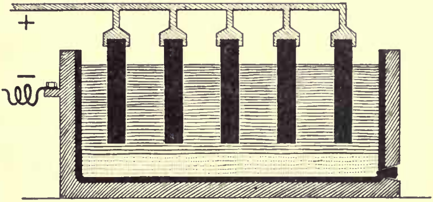


FIG. 55. — Manufacture of aluminium.

Properties of aluminium. Metallic aluminium has a

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[1] The plate by which the current enters is the *anode* and the one by which it leaves is the *cathode*.

[2] Cryolite is a white mineral found in Greenland. It is remarkable because a splinter of it will melt in the flame of a candle. It is sodium aluminium fluoride, Na_3AlF_6.

bluish white color, is light, having a specific gravity of only 2.6, and is quite soft. It is ductile, malleable, and tenacious. It is used as a conductor of electricity, for cups, saucepans, etc., and wherever strength and lightness are required. The metal keeps bright when exposed to the air. It oxidizes quite easily, but the thin film of oxide that forms is transparent and so does not show. This film protects the metal from any further action of the air.

Aluminium. It is satisfactory for cooking utensils. It is readily kept clean, and, as heat is easily transmitted through the thin metal, liquids are quickly heated in it. It has a decided advantage over enamel or agate ware in that it does not chip, wears much longer, and does not act as a heat insulator as does agate ware. It must be cleaned with a neutral substance, as strong acids and alkalies turn it black. It is not suitable for frying or pastry, as the metal heats so quickly that things stick and burn. It is ideal for teakettles and double boilers.

Aluminium is used in many alloys, as in aluminium bronze, an alloy of copper and aluminium having the color of gold and the strength of steel.

Aluminium oxidizes so easily that it is used in flashlights. The powdered metal is mixed with an oxidizing agent, when it burns with an exceedingly bright and actinic light.[1] As these flashlights are really a variety of explosive, they should be handled with care.

The powdered metal is also used as a paint. Celluloid is soluble in amyl acetate, and this solution, when mixed

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[1] Blue light aids chemical action much more than red light. It is for this reason that we use red lamps in a photographic dark room. We say that these blue rays of light are very *actinic*. An actinic light is one rich in the blue and violet rays that cause chemical action.

with aluminium powder, makes a good metallic paint for steam pipes and radiators.

Lead and its uses. Water pipes are sometimes made of lead. This soft, white metal melts at a low temperature, is easily bent, and is an ideal material for pipes, as it does not readily corrode in the air. All natural waters act somewhat on lead pipes, and, in some cases, the corrosion is so marked that it is not safe to use them. Lead hydroxide is formed, and the carbon dioxide present converts this into a carbonate. Lead salts are poisonous, and their action is cumulative. That is, lead is not eliminated from the body, and slowly accumulates until enough is present to produce illness. Painters often suffer from lead poisoning, as common paint contains white lead or basic lead carbonate.

Gold and its uses. Gold is a yellow metal. It is about as soft as lead, and is the most malleable of all the metals. By first rolling gold until it is thin, and then placing the thin sheets between gold beaters' skin and hammering, it can be beaten out so thin that 250,000 sheets would be only one inch thick. Dutch metal leaf, which is thin brass, is an imitation of gold leaf. The two can easily be distinguished, for the gold transmits a green light, while the Dutch metal is opaque.

Gold does not tarnish in air. For this reason, and because of its high cost, its beautiful color, and luster, it is used in jewelry. The pure metal would be too soft to use, as it would wear out very quickly. It is therefore alloyed with some harder metal. Silver and copper are used. Silver gives the gold a pale color, while copper makes it red. Other metals vary the color, arsenic, for instance, making it green.

The purity of gold is expressed in carats. Pure gold is 24 carats fine. A suitable fineness for general use is 14 carat, that is, fourteen pounds of gold to ten pounds of copper. This alloy is hard, of a good color, and is not too expensive. Gold coins contain nine parts of gold to one part of copper.

Articles should not be marked "solid gold." Suppose a pin is really eight carats fine. This means that it will contain two pounds of copper to one pound of gold, but, since the specific gravity of gold is more than twice that of copper, it means that, by volume, there will be only one sixth gold. If this pin is marked "solid gold," the purchaser imagines that he is getting pure gold, while in reality it is far from pure. It is much better to mark the exact carat, then the buyer knows what he is purchasing.

Silver and its uses. Silver is much harder than gold. The oxygen of pure air does not affect it, but air, especially the air of cities, contains hydrogen sulphide. This changes the silver to silver sulphide, and thus blackens it. The "oxidized silver" of the jeweler is really silver covered with black silver sulphide. To keep silver bright, it must be kept from this gas.

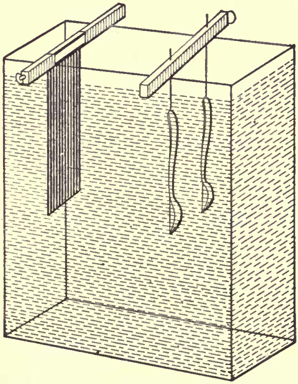


FIG. 56. — Silver plating.

Sterling silver is silver that contains, in each 1000 pounds, 925 pounds of silver. The remaining 75 pounds is metal used to alloy it, usually copper. Coin silver contains 90 % silver and 10 % copper.

Large silver teapots and other vessels are made by plating silver upon a white metal base. This is an alloy of lead and tin called pewter. The thickness of the coating is denoted by the name, as "triple plate." This does not mean that the vessel has been plated three times, but that three times the standard amount of silver has been deposited upon it. Silver is plated in the same way as nickel, using a plate of silver for the anode, and the article to be plated as the cathode (Fig. 56). The bath is usually silver cyanide, dissolved in potassium cyanide.

Copper. Pure copper is a rather soft, malleable, ductile, reddish metal, that is a very good conductor of both heat and electricity. It is used in large quantities in electric light and telephone wires, and in other electrical work. Its high thermal conductivity also makes it suitable for kitchen pots and pans. It corrodes easily, and as copper salts are poisonous, care must be used to keep kitchen utensils bright. It is easily cleaned, either with an abrasive, or by oxalic acid or ammonia, which dissolves the tarnish.

Use of alloys in sprinkler heads. Many of the metals when melted mix with each other to form solutions known as alloys. An alloy is usually more fusible than the metals that compose it. Thus Wood's metal is made up of tin and cadmium, each one part, lead two parts, and bismuth four parts. It melts at 165° F., which is lower than the melting point of any of its constituents. This low melting point of the alloy is made use of in *sprinkler heads* (Fig. 57). Many factories and stores are now protected against fire by covering the ceilings

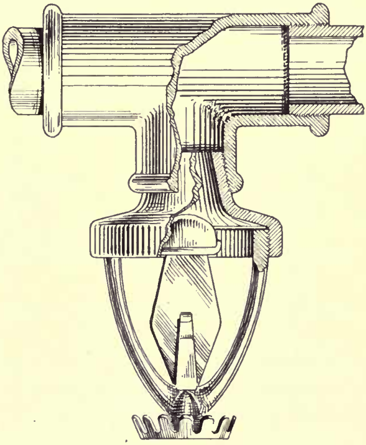


FIG. 57. — A sprinkler head.

with a network of pipes having at short intervals openings closed normally with a plug of Wood's metal. Should a fire occur when the building is unguarded, the heat will melt the easily fusible metal. Water will then flow through the pipes, and, falling upon the fire, extinguish it. Thus the damage is small. The system is usually so arranged that the melting of a plug not only turns on the water, but rings an alarm in the watchman's room.

Some household alloys. Some alloys are of particular interest in the household. Brass is an alloy of copper and zinc. Solder an alloy of lead and tin. The larger the percentage of lead, the more easily the solder melts, but also the weaker it is. The small strips of "soft solder" sold to mend tin kettles are of small value, as they contain so much lead that they have little strength. Bronze is an alloy of copper and tin. It usually also contains some zinc.

Amalgams. The alloys obtained by dissolving metals in mercury are given a special name, amalgams. Teeth are often filled with an amalgam. Gold is sometimes extracted from its ores by grinding the powdered wet ore with mercury. The mercury dissolves the gold, forming gold amalgam. On heating this, the mercury volatilizes, leaving the gold.

SUMMARY

A metal is a substance whose hydroxide is a base.

Iron is obtained by heating iron oxide, coke, and a flux in a blast furnace.

Galvanized iron is iron coated with zinc.

Tin plate is iron coated with tin.

Nickel is plated on iron, to protect the iron from rusting, as well as to give a good appearance.

Aluminium is obtained by electrolyzing a solution of aluminium oxide in fused cryolite.

Lead is poisonous, and care must be taken in using lead pipes.

Gold is the most malleable metal. Pure gold is 24 carats fine.

Coin silver is 90% silver.

Sterling silver is 92.5% silver.

Copper is a soft, malleable metal, used for electrical conductors and kitchen utensils. Its salts are poisonous.

Exercises

1. How could you harden a knife blade that is too soft?

2. Why is the soil in many localities red?

3. Why does a spot of rust often form on a tin pan?

4. Why are galvanized pails used instead of zinc pails?

5. Why is an aluminium teakettle better than an iron one?

6. Would you rather have a 22-carat, or a 14-carat gold ring? Why?

7. Aluminium is worth more per pound than brass, but aluminium rod costs less per foot than brass rod of the same size. Why?

8. How can you distinguish imitation gold leaf from the genuine?

9. In putting away your silver for the summer, how should you pack it, and why?

10. Why let the water run before using any for coffee or cereal, if you are preparing breakfast for the family?

11. Why should painters be especially careful to wash their hands before eating?

12. How can tin foil be sold for 30 cents a pound, while tin is worth 50 cents a pound?

CHAPTER XVI

PHOTOGRAPHY

Effect of light on substances. We see in our daily life many instances of the effects that light and air have upon various materials. The newspaper that to-day is white, will in the course of a few weeks turn to a yellow, and, in a few years, to a dull brown. The delicate colors of ribbons quickly fade in the hot sunlight of summer. Carpets fade in the sunlight, so that some too careful housewives pull down the shades to exclude the sun, and so deprive themselves of necessary light.

It is a simple matter to use these facts to obtain a photograph. If a leaf is laid upon a piece of blue wrapping paper, which is colored with a fugitive dye, and the paper then exposed to the light, the color will bleach except where the leaf protects it from the action of the light, and so we shall obtain a picture of the leaf.

The difficulty comes in preserving the picture thus obtained. In time, all of the blue dye will fade, and the picture disappear. Then, too, such pictures would be unsatisfactory because too long a time is required to produce them, and because the half tones are deficient.

Effect of light on silver compounds. The ordinary photograph depends upon the sensitiveness of silver salts to the action of light. When silver is dissolved in nitric acid, silver nitrate is obtained. This, when crystallized, is a white, heavy solid, $AgNO_3$. By careful heating it can be melted without decomposition, and cast into sticks. It is then called *lunar caustic*, and is used to burn away growths on the body, as it destroys flesh when brought in contact with it.

If you will examine the bottle in which the laboratory solution of silver nitrate is kept, you will notice that where the solution of silver has trickled down the outside of the bottle, it has left black streaks. The deposit is really metallic silver, but because it is so finely divided, it looks black, and you do not recognize it as silver.

Advantage is taken of this fact to make an indelible ink. If a solution of silver nitrate is mixed with a little gum, and then used as an ink to mark cloth, the marks will at first be colorless, but, on exposure to light, they gradually turn black, and an indelible mark is left.

We might use silver nitrate in photography, but the silver halogen compounds (silver chloride, bromide, and iodide) are better, as they are more sensitive to light, and are insoluble.

Principles of photography. When silver nitrate is mixed with a soluble chloride, as sodium chloride, a precipitate of insoluble silver chloride is formed.

$$AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$$

This silver chloride is a white insoluble compound that turns deep violet on exposure to light.

To make a better photograph of our leaf than we could make by bleaching our colored paper, we might first soak a piece of filter paper in salt, and then in a solution of silver nitrate. In this way silver chloride is precipitated right in the paper fibers. If now this paper is exposed to light with the leaf over it, the paper blackens except where protected from light by the leaf, and so a photograph of the leaf is obtained.

This process is also unsatisfactory, because, even though we keep the photograph in a dark drawer, it will eventually all turn black, and the picture disappear. A final step in the process is necessary to make the print permanent. Silver chloride is soluble in sodium thiosulphate, commonly called sodium hyposulphite, which, in turn, is shortened by photographers to "hypo." If then the print is soaked in hypo, the silver chloride that has not been acted on by the light will dissolve. No compound that is sensitive to light will then be left in the paper, that is, the print will be permanent.

The picture will, however, be reversed as regards light and shade. It will form what the photographer calls a negative.

In order to secure a correct reproduction of the original, both light and shade must be again reversed. This may be done by exposing a fresh piece of silver chloride paper under the negative. Wherever the negative is black (opaque), there the paper will be protected from the light, and so will remain colorless, while the translucent parts of the negative will transmit the light, and the paper under them will turn dark. Thus what is light in the negative will be dark in the copy. That is, a positive, or correct, reproduction of the leaf will be obtained.

Developing the negative. Silver bromide and iodide may be obtained by precipitating silver nitrate with a soluble bromide or iodide.

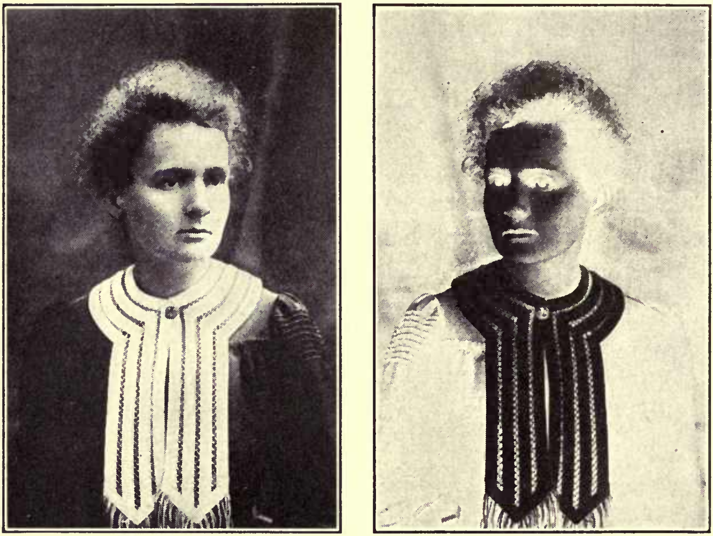
$$AgNO_3 + KBr \rightarrow AgBr \downarrow + KNO_3$$
$$AgNO_3 + NH_4I \rightarrow AgI \quad + NH_4NO_3$$

They are bodies much like silver chloride, but are even more sensitive to the action of light. Light does not discolor them as rapidly as it does the chloride, but a brief exposure to light causes a chemical change in them that, once started, can be continued by the use of what is known as a developer. The exact chemistry of the operation is very complex; it will be sufficient if you understand the following.

When light falls upon silver bromide or silver iodide in the presence of organic matter, as gelatin, it causes a change that is not well understood. The eye can see no change, yet when a developer (reducing agent) is added the affected silver compound is reduced to metallic silver. The silver bromide or iodide that has not been exposed to light is not reduced by the developer. In this way the *latent image*, as it is called, is rendered visible during development.

The photographic plate that you buy is made by slowly adding silver nitrate to the hot solution of a bromide and iodide in gelatin. This precipitates silver bromide and iodide as a fine powder. The character of the plate is determined by varying the proportion of the iodide to the bromide. The melted emulsion is then spread on glass plates. The gelatin cools, sets, and the plate is then dried. If a film is to be made, the same process is carried out, except that the emulsion is coated on a transparent film of celluloid. As the plates are sensitive to ordinary light, all of these operations must be carried on in a dim red light.

The photographic plate thus prepared is then exposed in the camera. Wherever light falls upon it, the silver compound is altered in such a way that when the plate is put into the developer, which is some reducing agent, the developer converts the altered silver compound to metallic silver. Since the chemical action of light is proportional to the amount of light that falls upon the plate, we obtain, on developing, a plate which gives a reproduction of the light and shade of the original, but a reproduction in which light and shade are reversed, that is, a negative (Fig. 58).



The positive. The negative.

FIG. 58. — Madame Curie, discoverer of radium. The light parts of the original are dark in the negative.

To illustrate: suppose we are photographing a person wearing a black dress with a white collar. The white collar will reflect much light to the plate. Where the image of this collar falls on the plate, there will be much chemical action, and on developing we shall have a black deposit representing the white collar. The dress, being black, will reflect only a little light to the plate; as a result the chemical action will be feeble, and on developing only a little silver will be reduced. The image of the black dress will therefore be almost transparent in the negative.

Fixing the negative. To render the negative permanent, it is then soaked in a solution of sodium thiosulphate, or hypo. This dissolves the silver salts that have not been acted upon by the developer, leaving nothing in the film but metallic silver. This silver is in such a fine state of division that, instead of looking bright, as you might expect, it is black. On washing the hypo out of the film and drying the plate, we have the finished negative. As this consists of silver particles embedded in gelatin, there is no reason to suppose that it will not last forever, if it is kept dry and is not broken.

Photographing colors. The amount of chemical action on silver salts produced by light depends on not only the amount of light, but also its color. The silver compounds are only slightly sensitive to red light, while they are very sensitive to blue light. This makes trouble in photographing. A white dress having both pale blue and pink dots on it will illustrate this. To the eye both pink and blue may seem to be equally bright. Since, however, silver salts are much more sensitive to blue than to pink, on photographing the dress the pink dots will reproduce almost black, while the pale blue will reproduce almost white. This causes the photograph of the dress to look unnatural. The same difficulty occurs in photographing flowers; some colors reproduce too light, others, too dark.

The difficulty may be overcome by adding certain dyes to the coating on the plate. These dyes make the plate sensitive to all the colors. Such plates are called *panchromatic*. They can be made to reproduce all colors satisfactorily, but are difficult to handle, as they are sensitive even to the red light of the dark room.

Photographic prints. An easy way to obtain a print showing light and shade as in the original is to expose under the negative a piece of paper having on its surface the same kind of coating that is on the ordinary plate, but so modified as to be less sensitive to light. On exposure to light, wherever the negative is opaque, as the image of a white collar, little light will pass through the negative to the paper and little chemical action will take place. Wherever the negative is transparent, as under the image of the black dress, much light will pass through, and there will be a large amount of chemical action on the silver salt. On developing, in the same way that the negative is developed, an image in metallic silver is obtained, but this time the light and shade are the reverse of those in the negative. This second reversal will bring objects to their natural shading. The white collar will be white and the black dress black.

The print must now be placed in hypo to fix it, that is, to dissolve the unaltered silver compounds, and on washing and drying will be permanent. "Velox" and "Cyko" are papers of this description.

Blue prints. Another common way of making positives is by the use of blue-print paper. The use of this paper depends upon two facts. First, ferric compounds, on exposure to light, especially in the presence of organic matter, are easily reduced to ferrous compounds. Second, potassium ferricyanide gives no precipitate with ferric compounds, but gives a bright blue precipitate with ferrous compounds.

A solution containing ferric ammonium citrate and potassium ferricyanide is prepared and coated on the surface of white paper. This is dried. If it is placed in water, the coating will dissolve, and leave nothing but a sheet of white paper. But let us place a sheet of this prepared paper under the negative that we have made. The image of the white collar being opaque will not allow light to pass, and there the coating will remain in its original condition. Under the image of the black dress, which is transparent, the light will act on the coating and reduce the ferric iron to ferrous. If now we put the paper into water, the ferrous iron and the ferricyanide will react, forming a blue precipitate, which will color the paper blue. We shall then have a picture (positive) in blue. Such paper is cheap, and easily prepared at home. It has the disadvantage of requiring a long exposure to light, and of giving a print that is somewhat lacking in half tones. The color, too, is not suitable for many subjects.

Printing-out papers. Still another kind of photographic paper is the so-called "printing-out paper" (P. O. P.), of which "Solio" is an example. It consists of paper coated with gelatin containing silver chloride. It is similar to blue-print paper in requiring a strong light for exposure, and in allowing the progress of the process of printing to be followed by inspecting the print from time to time. The image is formed directly by the action of the sunlight, no developer being required, therefore, to make it visible, as in the case of gaslight papers.

The paper is fixed by means of the usual hypo bath, followed by a final washing in water. But the image thus produced has a disagreeable brick-red color, so it is customary to give the image a more pleasing tone. This coloring or toning process, as it is called, is accomplished by placing the print after removal from the printing frame in a solution of chloride of gold. Here the silver particles are gold plated to any desired degree, the brick-red color changing first to a rich brown, and then to a purple tone. When the desired tone is reached, the print is fixed in hypo, then thoroughly washed, and dried.

SUMMARY

A photographic plate consists of a coating of gelatin containing silver salts on a transparent glass or celluloid support. On exposure to light and developing, the silver salt is reduced to metallic silver. On fixing in hypo, we obtain a *negative.*

A positive print can be made by exposing a piece of paper coated with a silver salt in gelatin under the negative, developing, and fixing.

Exercises

1. Why is it necessary to fix the negative?

2. Why do silver prints sometimes fade?

3. Why do you wash the negative after fixing?

4. Apples are sometimes sold that have the name of the grower in green letters on the skin, while the rest of the skin of the apple is red. Can you explain how these could be made?

5. Could you make a photograph using nothing but newspaper?

6. In photographing a landscape, why is it that faint clouds often do not show on the negative?

7. Why does the " proof " sent by the photographer fade?

CHAPTER XVII

CARBON AND ITS COMPOUNDS

Organic chemistry defined. One of the facts that every cook learns early in her experience is that all foods char when they are heated too hot. This is because all of these compounds contain carbon. Since all of these compounds that are made by living things come from organisms, we call them *organic*, and their chemistry, organic chemistry. We have learned to make many of these compounds in the laboratory. There is no real difference between the sugar made in the sugar cane and the sugar made in the laboratory, so we must call them both organic. Instead, then, of defining organic chemistry as the study of the chemistry of living organisms, it is preferable to say that organic chemistry is the chemistry of carbon compounds.

Forms of carbon. The element carbon is found widely distributed in nature in a number of forms, as the diamond, graphite, and coal. That these very dissimilar substances are all really carbon is proved by the fact that, when they burn, nothing but carbon dioxide is formed except in the case of coal, which gives some other products because of certain impurities in the coal. You will recall the fact that phosphorus occurs as both red and white phosphorus, two different forms having quite different properties, yet both forms can be shown to contain nothing but phosphorus atoms, and each can be readily converted into the other.

Chemists explain this by the fact that although both red and white phosphorus molecules contain nothing but phosphorus atoms, the number of atoms in the two molecules is different. Oxygen also exists in two forms, ordinary oxygen and ozone, having quite different physical and chemical properties. The same fact explains this. Ordinary oxygen is O_2, while ozone is O_3 Such different forms of the same element are called *allotropic modifications.*

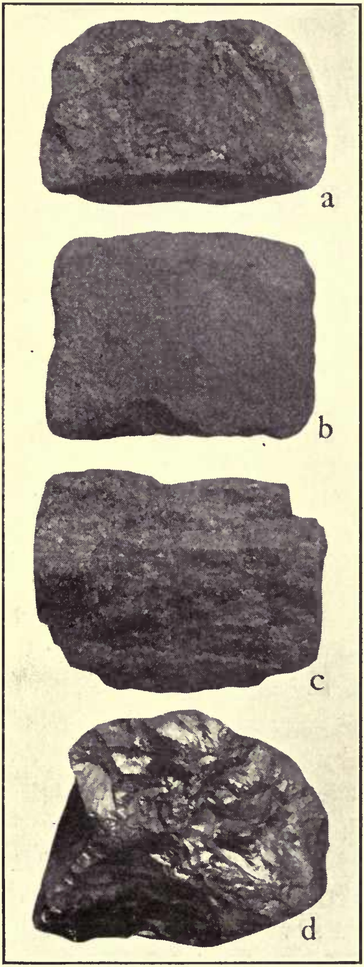


Fig. 59. — Coal series. *a*, peat; *b*, lignite; *c*, bituminous; *d*, anthracite.

Carbon is one of the elements that exist in a number of these allotropic modifications. It occurs in two main typical forms: crystalline carbon, including the diamond and graphite, and noncrystalline, or amorphous carbon, including lampblack and coal.

Formation of coal. Ages ago the distribution of land and water on the earth was very different from what it is to-day. Nor was the air of its present composition, as it contained more carbon dioxide and water vapor. By studying the rocks, and the fossils contained in them, geologists have given us a good idea of the conditions prevailing at that far-distant time. For instance, what is now the central part of Pennsylvania was then a low, swampy country. Its climate was much warmer than now, and, as the air contained large quantities of carbon dioxide and water, both of which are so important to growing plants, vegetation grew with almost tropical luxuriance. Ferns, for instance, were often thirty feet high and ten inches in diameter, as shown by their fossil remains. As this dense vegetation growing on low, swampy land died, it fell into the water. There it could not easily decay because of the lack of oxygen. Instead, as more and more accumulated, it turned into a solid mass similar to the peat of the present day. This accumulation of organic matter finally became buried deep beneath the surface of the earth, where pressure and heat made it still more compact. This formed what we now call brown coal, and large deposits of this exist all over the world. Brown coal is also called lignite.

As the heat and pressure were increased, still more of the volatile matter was driven from the brown coal, and soft, or bituminous, coal resulted. In a few places the process has been carried one step further, and hard, or anthracite, coal formed. In much of the soft coal we find the imprint of leaves and the remains of tree trunks, showing its vegetable origin. When soft coal is cut into very thin slices, it is even possible to see the cells of the wood from which it was made.

Wood to coal. Wood contains much volatile matter, and the progressive change of the wood (vegetable matter) into hard coal was essentially the driving off of this volatile matter, leaving only carbon. Hard or anthracite coal is, then, essentially carbon. It contains in addition the mineral matter of the original vegetation, and this is left as an ash when the coal is burned. These facts will explain to you why there is so much brown coal, less bituminous coal, and very little anthracite coal. Originally it was all brown coal. In a number of places the conditions were favorable, and the brown coal was changed into bituminous coal, while only in a few localities did the change continue, resulting in the formation of anthracite. This gradual change, resulting in an increase in the amount of carbon and a decrease in the amount of volatile matter, is well shown in the following table:[1]

FUEL	PER CENT OF (EXCLUDING ASH AND MOISTURE)				PER CENT OF ASH	CALORIFIC VALUE PER GRAM IN CALORIES
	C	H	O	N		
Wood	45	6	48	1	1.5	2700
Peat	60	6	32	2	5–20	3500
Brown coal . .	70	5	24	1	3–20	6000
Bituminous . .	82	5	12	1	1–15	8000
Anthracite . .	94	3	3	—	6	8000
Charcoal . . .	95	1.7	3.4	—	4	8080
Coke	96	0.7	2.5	1	3–11	7700

Illuminating gas. When soft coal is heated out of contact with the air, the volatile compounds in it are largely broken up into illuminating gas, coal tar, and ammonia, while most of the carbon remains in the retort in the form of a hard, porous mass called coke.

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[1] Inorganic Chemistry, Smith.

Commercially, this operation is carried out in long slender retorts. The gases produced are led into water, where most of the ammonia and much of the coal tar dissolves. The gases are then led into large pipes called condensers, where the remaining coal tar and liquid hydrocarbons condense. To remove the last traces of ammonia and coal tar, the gases are then passed through towers filled with coke, kept wet with water. Here the remainder of the ammonia and coal tar is removed. It is next passed over iron oxide, to remove any sulphur compounds present and illuminating gas remains. This is the old-fashioned process (Fig. 60). Some gas is still produced by it, but more is produced by another method, which will be explained in the next chapter (p. 192).

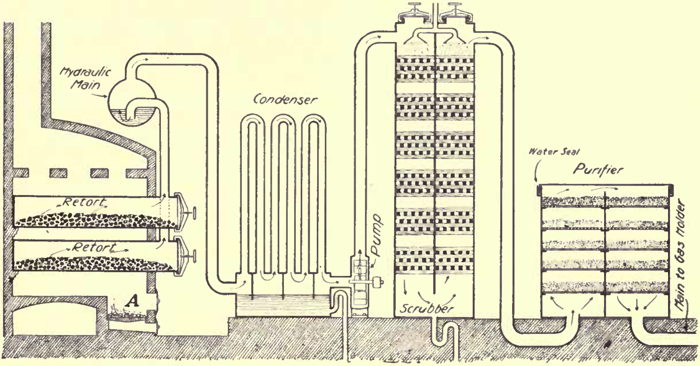

FIG. 60. — Coal gas manufacture.

The water solution of the ammonia is used commercially as a source of ammonia, while the coal tar contains many valuable carbon compounds, as benzol, used in the manufacture of aniline dyes, and naphthalene, used for moth balls. In Pennsylvania, soft coal is destructively distilled in large ovens for the sake of the coke, which is used largely in the production of pig iron in blast furnaces.

Petroleum. In many sections of our country, as in Pennsylvania, Texas, and California, far below the surface of the earth, large deposits of a thick, oily liquid exist. This liquid, called crude petroleum, is obtained by drilling wells (Fig. 61). Crude petroleum is not made up of one chemical compound, but contains a large number of compounds, and these are separated and prepared for use by distillation.

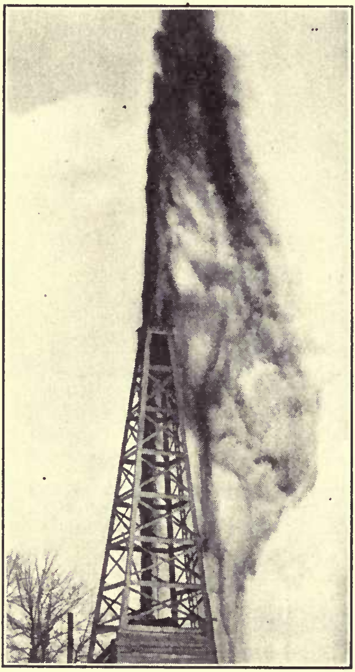

Fig. 61. — A spouting oil well.

The thick, oily liquid is placed in large stills, capable of holding 500 barrels or more of the oil (Fig. 62). Heat is then applied, and the more volatile liquids boil off first. The distillation products are called naphtha, benzine, or gasoline. After these have been driven off, the temperature rises and kerosene distills. Then come light, followed by heavy, lubricating oils, and finally nothing but pitch remains in the still.

Each of these products is a mixture of various carbon compounds. From the lubricating oil the waxy solids vaseline and paraffin are produced. The amount of petroleum produced is very large, as will be seen from the table below:

WORLD'S PRODUCTION OF PETROLEUM IN 1912

COUNTRY	QUANTITY. METRIC TONS	PERCENTAGE OF TOTAL
United States	29,615,096	63.25
Russia	9,317,700	19.37
Mexico	2,207,762	4.71
Dutch East Indies	1,478,132	3.09
Roumania	1,806,942	3.70
Galicia	1,187,007	2.43
India	989,801	2.03
Other countries	674,285	1.42
Total	47,276,725	100.00

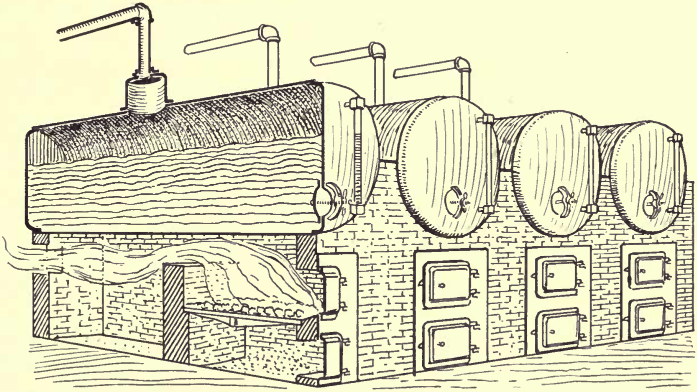

FIG. 62. — Petroleum stills.

Natural gas. We often find a combustible gas associated with crude petroleum. This gas, called natural gas, sometimes exists in such quantities that the oil wells furnish enough gas to warm and light whole cities. Pittsburgh is near an oil field that produced large quantities of this natural gas, and for years, instead of using artificial illuminating gas, the city used this natural gas. Burned in ordinary gas burners it does not give a bright light, but is well suited for use with Welsbach burners, and in gas stoves.

Charcoal and bone black. When wood is heated out of contact with the air (destructively distilled), the volatile matters are driven off. Many products are thus obtained, as wood alcohol, acetic acid, and combustible gases, while charcoal is left in the retort. Not only wood, but any organic matter can be destructively distilled, and a charcoal produced.

If bones are thus treated, bone charcoal, called *bone black,* is obtained. This has the power of absorbing many times its own volume of gas. If this bone charcoal is exposed to the air, it absorbs much oxygen, and because of this it can be used to oxidize noxious odors. It is also used in the manufacture of sugar to remove the brown color of the sugar sirup. It is an interesting experiment to shake together bone charcoal (bone black) and a solution of hydrogen sulphide, the substance that gives rotten eggs their characteristic odor. The odor is quickly destroyed, showing how energetic the action of the bone black is. The best way to prepare pure carbon is to char sugar, thus obtaining sugar charcoal.

Lampblack. Lampblack is made by the incomplete combustion of materials rich in carbon. Natural gas, for instance, is burned in such a way that the flame strikes an iron surface kept cool by water flowing over it. Under these conditions, the compounds present in the gas decompose, forming carbon and hydrogen. The hydrogen burns, forming water, but the carbon that forms in the flame is cooled below its kindling point before it can burn, by the cold iron against which it strikes, and is deposited on the iron.

A scraper removes the carbon from the iron, and it forms the lampblack of trade. It is largely carbon, but contains some tarry matters. It finds many uses, as in black paints, in printer's ink, and in darkening the color of cement sidewalks.

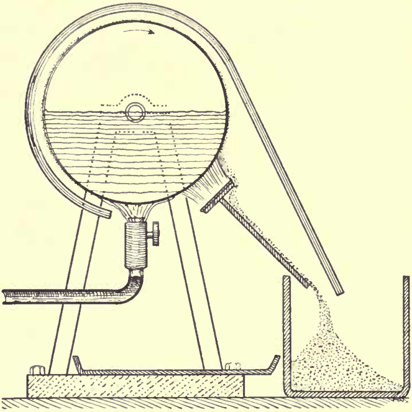

FIG. 63. — Commercial production of lampblack.

Graphite. Our "lead pencils" are made from another allotropic form of carbon, called graphite, plumbago, or black lead. We are not sure just how it has been formed in nature, but some chemists think that it has been formed by carrying the heating and pressure of coal a step beyond the point where anthracite is produced.

Graphite occurs as grains in rock, and is crystalline, forming six-sided plates. It is a black, greasy-feeling substance, having a metallic luster. It is used in making crucibles, as a lubricant, and in lead pencils. In spite of its seeming softness, the individual particles of graphite are very hard, as is shown by the way the iron tools used in cutting it wear out. Because of its greasy nature, it clings readily to metals, and is therefore used as a protective coating. Common stove polish, for instance, is largely graphite.

The various grades of hardness of lead pencils are obtained by using various-sized particles of graphite, and binding them together with clay. Large particles and little clay gives a very soft pencil.

Chemists have found it possible, by the help of the electric furnace, to prepare graphite artificially, and large quantities are now made at Niagara Falls, where electric power is cheap. This synthetic graphite is used in making the electrodes for electric furnaces.

Diamond. The diamond occurs in Brazil, India, and South Africa. It is valuable as a jewel because of its rarity, its hardness (it is the hardest substance known), its luster, and its effect on light. You have all looked through a glass prism and seen that rays of light are bent, on passing through it, and that the edges of objects seen through it are fringed with colored light. The diamond will bend and disperse white light much more than glass will, and it is this high refraction that gives the fire and color to the cut stone. A piece of glass, cut in the same form as a diamond, may be equally transparent, but the play of color will be lacking, and the beauty of the diamond is not found in the imitation.

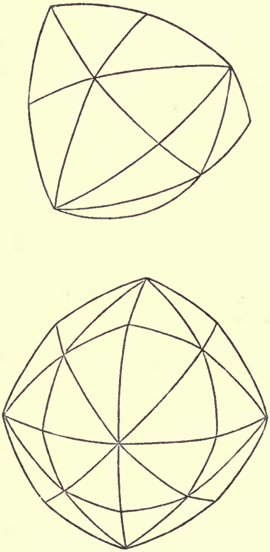

Fig. 64. — Crystal forms of natural diamonds.

When first found, diamonds are often shaped like an octahedron (Fig. 64). They have little beauty, and must be cut in a regular form in order to bring out their hidden splendor. The stone is so hard that it cannot be ground with any of the usual abrasives. It must be cut and polished by the use of diamond dust.

The diamond that is to be cut is first set into a lump of solder, fastened to the end of a short stick. The part of the diamond that is to be ground off is then laid on an iron plate that can be rotated, and the plate smeared with diamond dust mixed with oil. The iron plate is rotated and the diamond is very slowly worn away by rubbing on the diamond dust. When one face of the stone has been thus polished, the solder is heated, the diamond taken out, and put back so as to expose another face that is to be polished. In this way, face after face is polished, until the stone has its finished form. This takes a long time, requires skilled labor, and so adds much to the cost. Diamond dust is made by crushing small imperfect stones, for, although the diamond is

Fig. 65. — Sorting diamonds.

so hard, it is also very brittle. It is not an unusual thing for a diamond to break if dropped on a stone floor.

The value of a diamond depends on its freedom from flaws, the perfection of the cut, the color, and its fire. Diamonds of many colors are found, pink, blue, brown, and yellow, as well as white. A stone that has a decided color is often more valuable than a white stone, but a faint tinge of yellow kills the fire of the stone, and decreases its value greatly.

The largest diamond that has been found was, before cutting, the size of a man's fist, and weighed 621 grams, or more than one pound. The famous Kohinoor, one of the crown jewels of England, weighs 22 grams (Fig. 66). In Brazil, black diamonds are found. These are of no value as jewels, but, owing to their great hardness, they are used to point drills for drilling hard rock.

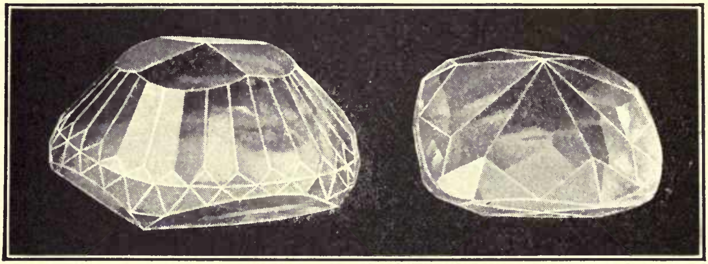

Fig. 66. — Kohinoor diamond. Weight of first cut, 186.5 carats; recut, 125 carats.

The weight of a diamond is always given in carats. The new international carat weighs 200 milligrams. The value of a diamond increases more rapidly than its weight. If a first-class cut stone of $\frac{1}{2}$ carat is worth 100 dollars, a stone weighing one carat will be worth 275 dollars, and one of two carats' weight will be worth 650 dollars. The cost of a one-carat diamond may be anywhere between one hundred and three hundred dollars, depending on the color, perfection of cut, and fire of the stone.

Small diamonds have been made synthetically, but they are only chemical curiosities, as they are costly, and very small.

SUMMARY

Carbon has three allotropic forms: diamond, graphite, lampblack.

Coal was formed by the partial oxidation of vegetable matter out of contact with the air.

Bone black is an important form of carbon, much used in decolorizing impure solutions.

Sugar charcoal is a pure form of carbon easily prepared.

Crude petroleum yields gasoline, benzine, naphtha, kerosene, lubricating oils, vaseline, and paraffin.

Graphite is used in lead pencils, stove polish, crucibles, and electrodes for electric furnaces.

The diamond is the hardest substance known.

Exercises

1. How can you prove that graphite is carbon?

2. Why does a soft-coal fire give a much brighter light than a hard-coal fire?

3. Why is charcoal sometimes placed in a refrigerator?

4. Why does bone black lose its efficacy after long use?

5. Is coal being made at the present time?

CHAPTER XVIII

THE OXIDES OF CARBON

Preparation and properties of carbon dioxide. Whenever a carbon compound burns with a free supply of air, a colorless gas, carbon dioxide, CO_2, is formed. The pure gas is prepared by the action of an acid on a carbonate. Place marble chips (calcium carbonate) in a generator, similar to the one used in preparing hydrogen (Fig. 67). Add hydrochloric acid through the thistle tube. A violent evolution of gas takes place, and this gas can be collected by water displacement.

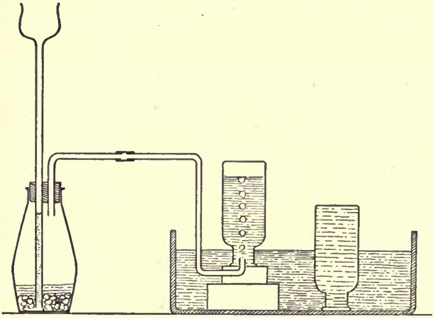

FIG. 67. — Carbon dioxide generator.

$$CaCO_3 + 2\,HCl \rightarrow CaCl_2 + H_2O + CO_2\uparrow$$

Any carbonate may be used instead of marble, and any acid that will dissolve the carbonate, as:

$$Na_2CO_3 + H_2SO_4 \rightarrow Na_2SO_4 + H_2O + CO_2\uparrow$$

Examination of the carbon dioxide thus prepared shows that it is a colorless gas, without taste or odor. When tested with a lighted match, it does not burn or support combustion. On shaking it up with a little water, it dissolves, giving a very slightly sour solution that turns blue litmus red. Since the solution of carbon dioxide in water gives an acid reaction with litmus, it must in solution form an acid.

$$H_2O + CO_2 \rightarrow H_2CO_3$$

We name this acid carbonic acid, and the salts made from it carbonates.

One thing that distinguishes carbon dioxide from other colorless gases is its specific gravity. It is about one and a half times as heavy as air. This can be shown in a striking way by placing a candle in a glass, lighting it, and then pouring carbon dioxide down on it from a bottle. Carbon dioxide is so heavy that it can be poured like water; it therefore fills the glass and extinguishes the candle.

Carbon dioxide gas is often incorrectly called carbonic acid. Of course it cannot be an acid, as all acids must contain hydrogen. It is only when it is dissolved in water that it becomes an acid. It may properly be called carbonic acid anhydride.

FIG. 68. — Carbonated spring waters often deposit dissolved solids, making a raised cone from which the water flows.

Uses of carbon dioxide. The great uses of carbon dioxide are in beverages and as a leavening agent. Every drink that effervesces, soda, vichy, root beer, etc., owes its effervescence to the presence of dissolved carbon dioxide. When you drink plain soda, you are really drinking a weak solution of carbonic acid.

Its use as a leavening agent will be treated in the following chapter (p. 196).

In many parts of the earth, spring waters naturally contain considerable amounts of this gas. The springs at Saratoga are illustrations of this, as are the Vichy, Seltzer, and Apollinaris springs.

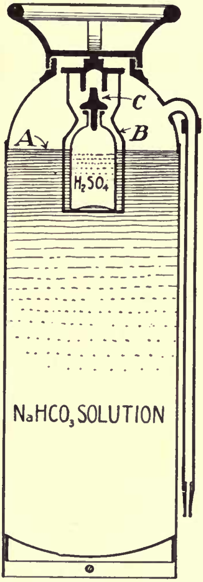

FIG. 69. — Section of a fire extinguisher.

Fire extinguisher. As carbon dioxide will neither burn nor support burning, it is used in fire extinguishers. In many buildings you have seen copper cylinders with a short hose coming from the top. Examine one and you will find the directions: "To start, turn upside down, and play on the fire." Their construction is simple. The cylinder is filled with a solution of sodium bicarbonate. At the top is suspended a bottle containing sulphuric acid. This bottle is loosely closed with a lead stopper. On turning the cylinder upside down, the stopper falls out and the acid runs out. When the acid mixes with the sodium bicarbonate, carbon dioxide is generated. Some of this dissolves in the water, and the rest, accumulating in the cylinder, generates a pressure that forces the water and gas through the hose upon the fire. They are efficient if used before the fire has too great a start.

$$2\,NaHCO_3 + H_2SO_4 \rightarrow Na_2SO_4 + 2\,H_2O + 2\,CO_2\uparrow$$

Test for carbon dioxide. It is easy to test for the presence of carbon dioxide by the use of a solution of calcium hydroxide, called limewater. If a little limewater is shaken with carbon dioxide, calcium carbonate is formed, and the solution turns milky.

$$Ca(OH)_2 + CO_2 \rightarrow CaCO_3 \downarrow + H_2O$$

Utilization of carbon dioxide by plants. Every fire produces large amounts of carbon dioxide, animals breathe it out; it is one of the products of the decay of all organic material. In this way large amounts get into the air. Plants then utilize it as a food. Under the influence of the energy of sunlight, chlorophyll, the green coloring matter of plants, makes starch from it. Much of the oxygen of the carbon dioxide is not needed in this process, and the plant passes this once more into the air. In this way, the amount of oxygen in the air is kept constant.

Carbon monoxide. When carbon is burned in an amount of air too small for its complete combustion, a second oxide of carbon, carbon monoxide, CO, is formed. It is a colorless, tasteless, odorless gas, very slightly soluble in water. It burns with a peculiar pale blue flame, forming carbon dioxide.

$$2\ CO + O_2 \rightarrow 2\ CO_2$$

Carbon monoxide is a poisonous gas, and is all the more dangerous because, owing to its lack of odor, we have no warning of its presence. Water gas, which is the illuminating gas commonly in use, contains about 40 % of it, and it is the presence of this carbon monoxide that makes gas leaks so dangerous.

The oxygen necessary to oxidize our tissues is carried to the remotest parts of our body by the red coloring matter of the blood, called hæmoglobin. This combines in the lungs with oxygen, forming with it a loose chemical compound. The blood then carries the hæmoglobin, combined with oxygen, wherever it is needed. When this oxygen is given up to oxidize materials in our body, carbon dioxide is produced as a waste product. The blood carries this carbon dioxide to the lungs, where it is exhaled, and where the hæmoglobin combines with more oxygen. Carbon monoxide forms a stable compound with hæmoglobin. This makes the combination of the hæmoglobin with oxygen impossible, the body cannot obtain oxygen, and we suffocate.

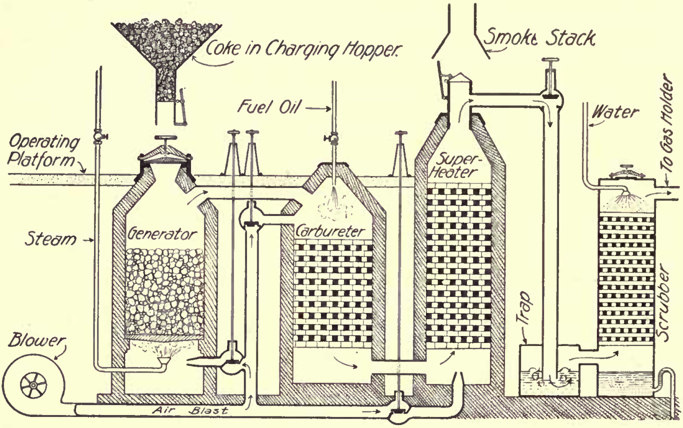

FIG. 70. — Manufacture of water gas.

The taking up of oxygen by the hæmoglobin gives the blood its bright red color, while its combination with carbon dioxide gives the blood a purple color.

Water gas. There are two ways of making illuminating gas. The first, by heating soft coal (p. 178). The second is by the action of steam upon red-hot carbon. Coke is placed in large iron cylinders and air blown through it until the coke is white hot. The air blast is then shut off, and steam blown through the coke. The white-hot coke reduces the steam, forming carbon monoxide and hydrogen.

$$H_2O + C \rightarrow H_2 + CO$$

When the steam has cooled the coke so that the reaction is slow, air is once more blown through the coke, it is again heated white hot, and the process is repeated.

This mixture of hydrogen and carbon monoxide burns with a hot but colorless flame. It is suitable for use in a gas stove or a Welsbach burner, but not for use in an ordinary gas burner for lighting purposes. To make it burn with a luminous flame, it must be mixed with the vapor of an oil that will decompose in the flame and set carbon free. To do this, the gas is passed into the carburetor, where it is mixed with a spray of crude oil. It then passes to a very hot chamber called a superheater, where the mixture is made complete and the oil vapors are changed into permanent gases.

The water gas must then be purified by passing it into a wash box, where it is cooled by water, and some tarry products that have been formed are condensed. Then it goes to the scrubber, a tower where the gas has to rise through zigzag passages that are kept wet. Finally, it passes to the condenser pipes, where it is cooled to 150° F., and finally to the gas holder.

The candle power of the gas is controlled by the amount of oil that is added to it. The process is cheap, and the product burns well and gives a good light. Its disadvantage is the poisonous character of the gas, due to the large percentage of carbon monoxide that it contains. On the average more than one person a day dies in the city of New York through inhaling water gas.

Danger from furnace gas. Carbon monoxide is sometimes made accidentally in ordinary furnaces. If the bed of coal is quite thick, and the amount of air somewhat limited, the oxygen of the air will all be used up in the lower part of the fire in combining with the burning carbon and forming carbon dioxide. This carbon dioxide then passes through the upper part of the fire, where the hot carbon reduces it to carbon monoxide. The carbon monoxide thus produced burns on the top of the fire and is the cause of the flickering blue flame sometimes seen playing over the top of an anthracite fire.

SUMMARY

Carbon dioxide: *How produced.* Carbon dioxide is prepared by the action of an acid on a carbonate. It is formed by the decay of organic matter, the combustion of fuels, and fermentation.

Properties: Carbon dioxide is a heavy, colorless, tasteless, odorless gas. It does not burn or support combustion. It dissolves in water, forming carbonic acid.

Uses: Carbon dioxide is used in beverages, as a leavening agent, in fire extinguishers, and as food for plants.

Test: Limewater is a test for carbon dioxide.

Carbon monoxide: *How produced.* Carbon monoxide is made by heating carbon dioxide with carbon.

Properties: Carbon monoxide is a colorless, tasteless, odorless gas. It is a poison.

Water gas is a mixture of carbon monoxide, hydrogen, and gases from crude oil. It burns with a luminous flame, and is very poisonous.

Exercises

1. The city of New York burns enough coal to convert all of the oxygen of the air over the city into carbon dioxide. Why do the inhabitants not suffocate?

2. When you open a bottle of vichy, the contents of the bottle become milky with gas bubbles. Why?

3. Why can we not obtain pure carbon dioxide by burning charcoal in the air?

4. Why is carbon dioxide used in fire extinguishers?

5. Why is water gas enriched with oil vapors?

6. Under what conditions may a poisonous gas be given off from the kitchen range?

CHAPTER XIX

BAKING POWDERS

Mechanical methods of leavening bread. Bread, our great staple food, and cakes and pastries with which we tempt our palates, are usually made light, to render them more palatable and digestible. Bread is usually leavened by the use of yeast (p. 287). At present we shall consider some other methods of gaining the same end.

Unleavened bread, as used in hard-tack, the cornmeal " pone " of the South, and the Scotch oatmeal cake, contains few gas bubbles. In consequence, it is hard, dry, and demands long mastication. To render bread light, the dough must contain a multitude of small gas bubbles. When the dough is placed in the oven, these expand, the dough increases in volume, and the bread is made light.

There are a number of mechanical means by which this can be done. By vigorously beating the mixture of flour and water, air bubbles can be entrapped, and these will make the bread somewhat light. A thin dough must be used, as otherwise the beating is too difficult. The addition of eggs that have been beaten to a froth will add air bubbles. This is the method used in making sponge cake. The water used in mixing may be replaced by soda water. This will cause lightness by the presence of carbon dioxide bubbles. This, however, is not a suitable method for the home, since the mixing must be carried out in air-tight vessels, else the gas will escape.

Chemical methods of leavening bread. All of these mechanical methods, while they may be used, do not satisfy the home requirements of a quick, easy method of leavening. For the home we must resort to some chemical process. Recall the method of preparing carbon dioxide by mixing a carbonate and an acid. If we can find two such compounds which can be mixed with the flour, and which will not give any disagreeable by-products, our problem is solved. There are many such combinations possible. The old-fashioned "soda biscuit" will serve as an example. These biscuits are made by mixing sour milk, which contains lactic acid, with sodium bicarbonate (saleratus).

$$NaHCO_3 + HC_2H_5OCO_2 \rightarrow NaC_2H_5OCO_2 + H_2O + CO_2$$

This leaves nothing but sodium lactate in the biscuit, and there is no objection to this. There are, however, two possible causes for failure, the addition of too much or too little baking soda. The acidity of the sour milk will vary within wide limits, and so no receipt can tell just how much baking soda to use to neutralize it. If the milk is unusually sour, acid will be left and the biscuit will be sour. If the milk is not as acid as usual, saleratus will be left over. Heat will convert this into washing soda, and the biscuit will have yellow streaks and taste soapy in places.

Commercial baking powders. This difficulty of measuring out the exact amounts of acid and carbonate needed for neutralization applies to any home-made mixture, hence the wide use of commercial baking powders, where the necessary proportions have been determined by a chemist. There are three important varieties of baking powders, — cream of tartar, phosphate, and alum powders.

Tartrate baking powder. The acid taste of grapes is partly due to the presence of an acid tartrate. When grapes are pressed, and the juice allowed to stand, this acid tartrate is deposited as a dark pink crust on the sides of the vat. It is called argols, and, when purified by solution and recrystallization, yields pure potassium acid tartrate, $KHC_4H_4O_6$. This is often called cream of tartar, and is the acid used in cream of tartar powders. The carbonate used is sodium bicarbonate, $NaHCO_3$, called baking soda. These two, when mixed in the proper proportions, give an excellent baking powder.

$$\begin{array}{ccccccccc} KHC_4H_4O_6 & + & NaHCO_3 & \rightarrow & KNaC_4H_4O_6 & + & H_2O & + & CO_2 \\ 188 & + & 84 & \rightarrow & 210 & + & 18 & + & 44 \end{array}$$

If the mixed powder is placed in a can and kept for some time, it is difficult to avoid getting a little moisture in the can from the air. The presence of this moisture causes the acid and the carbonate to act on each other, and the powder spoils. To avoid this, 20% of starch is added. The starch coats each particle of the acid and carbonate and prevents their acting on each other. At the same time it only adds a little more flour to the dough when the powder is used, and so is not objectionable.

An unscrupulous manufacturer may, however, use not 20%, but 50% or more and so make a very inferior powder, and still one that he can advertise as perfectly pure. The amount of starch can be readily determined by stirring a teaspoonful of the baking powder in a glass of water. All of the materials are soluble except the starch, which will sink to the bottom of the tumbler. A comparison of two powders in this way will often show that a powder that is cheaper by the pound is really more expensive than a high-priced powder, owing to the large amount of starch that it contains.

Occasionally a small amount of ammonium carbonate is added. This volatilizes in the heat of the oven, and the ammonia is driven off, while the carbon dioxide produced helps in making the bread light.

The by-product of cream of tartar baking powder, potassium sodium tartrate or Rochelle salts, is left in the bread. This is a laxative. It is the same product that is formed by seidlitz powders. If large quantities of baking powder are used to make a cake unusually light, and we eat an extra piece of the cake because it is so good, the amount of Rochelle salts that we take may be more than is desirable. As a rule, the quantity is too slight to be objectionable.

Phosphate baking powder. The phosphate powders use baking soda and calcium hydrogen phosphate (calcium superphosphate).

$$\begin{array}{ccccccccccc} CaH_4(PO_4)_2 & + & 2\,NaHCO_3 & \rightarrow & CaHPO_4 & + & Na_2HPO_4 & + & 2\,CO_2 & + & 2\,H_2O \\ 234 & + & 168 & \rightarrow & 136 & + & 142 & + & 88 & + & 36 \end{array}$$

They require a filler, just as do all baking powders. The substances left by the reaction are not injurious.

Alum baking powder. The alum powders contain baking soda and alum, usually ammonium alum, as that is the cheapest. It is a question whether the residue left from alum powders is objectionable or not. Many doctors say that it is decidedly injurious, others that it is harmless. So long as there is a question about it, and there is nothing to be gained by their use, except a small saving in the cost, it is well to avoid them.

$$\begin{array}{ccccccccccc} (NH_4)_2Al_2(SO_4)_4 & + & 6\,NaHCO_3 & \rightarrow & 2\,Al(OH)_3 & + & 3\,Na_2SO_4 & + & (NH_4)_2SO_4 & + & 6\,CO_2 \\ 475 & + & 504 & \rightarrow & 157 & + & 426 & + & 132 & + & 264 \end{array}$$

Commercial baking powders are frequently mixtures of phosphate and alum powders.

Ratio of ingredients used in making baking powder. The manufacturer of baking powder knows that, if he puts cream of tartar and baking soda together, he will make baking powder. The question is, how much of each to use, and for this information he must turn to the chemist. The chemist knows that one molecule of baking soda will combine with one molecule of cream of tartar, and that, as a result, Rochelle salts, water, and carbon dioxide will be formed.

This fact he may express by the following equation :

$$\underset{\text{baking soda}}{NaHCO_3} + \underset{\text{cream of tartar}}{KHC_4H_4O_6} \rightarrow \underset{\text{Rochelle salts}}{KNaC_4H_4O_6} + \underset{\text{water}}{H_2O} + \underset{\text{carbon dioxide}}{CO_2}$$

Knowing the atomic weights, he can calculate the molecular weights by the following operation :

$$\begin{array}{ccccccccc} NaHCO_3 & + & KHC_4H_4O_6 & \rightarrow & KNaC_4H_4O_6 & + & H_2O & + & CO_2 \\ 84 & + & 188 & \rightarrow & 210 & + & 18 & + & 44 \end{array}$$

Sodium Bicarbonate

$$\begin{array}{lr} Na = & 23 \\ H = & 1 \\ C = & 12 \\ 3\,O = 3 \times 16 = & \underline{48} \\ \text{Mol. wt. of } NaHCO_3 & 84 \end{array}$$

Cream of Tartar

$$\begin{array}{lr} K = & 39 \\ H = & 1 \\ 4\,C = 4 \times 12 = & 48 \\ 4\,H = 4 \times 1 = & 4 \\ 6\,O = 6 \times 16 = & \underline{96} \\ \text{Mol. wt. of } KHC_4H_4O_6 & 188 \end{array}$$

Rochelle Salts

$$\begin{array}{lr} K = & 39 \\ Na = & 23 \\ 4\,C = 4 \times 12 = & 48 \\ 4\,H = 4 \times 1 = & 4 \\ 6\,O = 6 \times 16 = & \underline{96} \\ \text{Mol. wt. of } KNaC_4H_4O_6 & 210 \end{array}$$

Water

$$\begin{array}{lr} 2\,H = 2 \times 1 = & 2 \\ O = & \underline{16} \\ \text{Mol. wt. of } H_2O & 18 \end{array}$$

Carbon Dioxide

$$\begin{array}{lr} C = & 12 \\ 2\,O = 2 \times 16 = & \underline{32} \\ \text{Mol. wt. of } CO_2 & 44 \end{array}$$

Total weight used $\overline{272}$ = Total weight obtained $\overline{272}$

It is apparent that, if the manufacturer mixes 84 pounds of sodium bicarbonate with 188 pounds of potassium acid tartrate, he will have 272 pounds of baking powder that when used will leave an excess of neither ingredient. By possessing this exact information he not only prevents the waste of soda or cream of tartar, but produces a better baking powder. He also knows that 84 pounds of baking soda (sodium acid carbonate, or sodium bicarbonate) will combine with 188 pounds of cream of tartar (potassium hydrogen tartrate, or potassium acid tartrate), to produce 210 pounds of Rochelle salts (potassium sodium tartrate), 18 pounds of water, and 44 pounds of carbon dioxide. A simple proportion will then give the amount of each ingredient needed to produce any required amount of baking powder.[1] Suppose we wish to manufacture 100 pounds of baking powder. Since 84 pounds of baking soda will give 272 pounds of baking powder, x pounds of baking soda will give 100 pounds of baking powder. Or: $84 : 272 :: x : 100$. $x = 30.8^+$ pounds; and since 188 pounds of cream of tartar are needed to produce 272 pounds of baking powder, then x pounds of cream of tartar will produce 100 pounds of baking powder. Or: $188 : 272 :: x : 100$. $x = 69.1^+$ pounds. It will therefore take 30.8 pounds of baking soda and 69.1 pounds of cream of tartar to make 100 pounds of baking powder.

A home-made baking powder. A satisfactory baking powder can easily be made at home from the following:

Cream of tartar, dried . . .	1 pound
Cornstarch, dried	½ pound
Baking soda	½ pound

---

[1] Baking powder usually contains some filler, as starch. This, for the sake of simplicity, is omitted in the problem, as it takes no part in the chemical change.

Divide the cornstarch into two equal parts. Mix one part with the cream of tartar, and the other part with the baking soda. Then mix the two together, place in cans and in a dry place. The important things to remember are that the powders must be dry, and that they must be well mixed. The cream of tartar and the cornstarch may be dried in a warm oven. The baking soda, however, must be used as purchased, as heating it will drive off some carbon dioxide and convert it into washing soda.

SUMMARY

All baking powders liberate carbon dioxide when mixed with the dough. It is this gas that makes the bread light.

Cream of tartar baking powders contain potassium hydrogen tartrate, sodium bicarbonate, and starch.

Phosphate baking powders contain calcium superphosphate, sodium bicarbonate, and starch.

Alum baking powders contain alum, sodium bicarbonate, and starch.

Exercises

1. Sour milk and baking soda will set carbon dioxide free. Why buy expensive baking powders to do the same thing?

2. A mixture of marble and hydrochloric acid will set carbon dioxide free. Why not use this mixture instead of baking powder?

3. Why is baking soda instead of washing soda used in baking powders?

4. Could carbonated water be used to make bread light?

5. Is a baking powder at 40 cents a pound always cheaper than one at 50 cents a pound? Explain.

CHAPTER XX

HYDROCARBONS AND DERIVED COMPOUNDS

Hydrocarbons. One thing that simplifies the study of the compounds of carbon is the fact that many of them can be grouped into series. Thus, hydrogen and carbon unite in different proportions and form several hundred compounds. Many of these occur in nature, while many are the products of the laboratory. They are all called *hydrocarbons*. A study of a few of them will show us how they may be arranged in a series, and how knowing the properties of a few of the members of this series will enable us easily to remember the properties of all.

The paraffin series. When rowing, you have doubtless pushed your oar into the mud on the bottom of some pond, and noticed that bubbles of gas arose. This gas, methane, CH_4, is called marsh gas, and is formed when vegetable matter decays under water. It is the main constituent of natural gas.

In crude petroleum we find three other gases: ethane, C_2H_6; propane, C_3H_8; and butane, C_4H_{10}. On examining these four formulas, you will see that, arranging them according to the number of carbon atoms they contain, each gas differs from the next by the group CH_2. We may then express all of these formulas in one general formula, C_nH_{2n+2}. If, then, we wish to know the formula of a hydrocarbon containing 16 carbon atoms, we can find it by multiplying 16 by

2 and then adding 2. The compound must then contain 34 hydrogen atoms, and its formula is $C_{16}H_{34}$.

These hydrocarbons are named, above the first four, by using as a prefix the Greek numeral that tells us the number of carbon atoms present, and following it by the ending *-ane*. Thus, C_5H_{12} is called pentane. Reversing the process, if the name of the hydrocarbon is octane, we know that it is a member of this series because the name ends in *-ane*. It must contain eight carbon atoms because of the prefix meaning eight, and its formula must be C_8H_{18}. This series is known as the paraffin series, and its members have been prepared up to $C_{60}H_{122}$.

The properties of all of the members of the paraffin series vary in a regular manner. Methane is a gas. With difficulty it can be changed into a liquid having a boiling point of − 164° C. Ethane boils at − 89.5° C., propane at − 37° C., butane at + 1° C., pentane at + 35° C. Without studying hexane we can then be sure that it will boil at about 70° C. Other properties vary in a similar manner, so that the study of a few members enables us to predict the properties of all.

The ethylene and acetylene series. There are several such series of hydrocarbons. One starts with the gas ethylene, C_2H_4. This is the gas that is present in illuminating gas, and makes the flame luminous. The general formula of this series is C_nH_{2n}.

A third series starts with acetylene, C_2H_2. It is easily made by the action of water on calcium carbide:

$$2\,H_2O + CaC_2 \rightarrow C_2H_2 + Ca(OH)_2$$

As calcium carbide can be cheaply made in the electric furnace from carbon and lime, acetylene can be cheaply prepared.

$$CaO + 3\,C \rightarrow CO + CaC_2$$

Acetylene is used largely in the lamps of automobiles, and in lighting country homes where ordinary illuminating gas is not available. Automatic generators are used, so arranged that the gas is made only as it is burned (Fig. 71). They require little care, and give acetylene, which burns with an exceedingly bright white light. The combustion of one cubic foot of acetylene gives fifteen times as much light as the combustion of one cubic foot of illuminating gas. For this reason a special burner is required which permits only a very small jet of gas to escape to feed the flame (Fig. 72). An oxygen acetylene flame gives an intense heat and is used in welding metals.

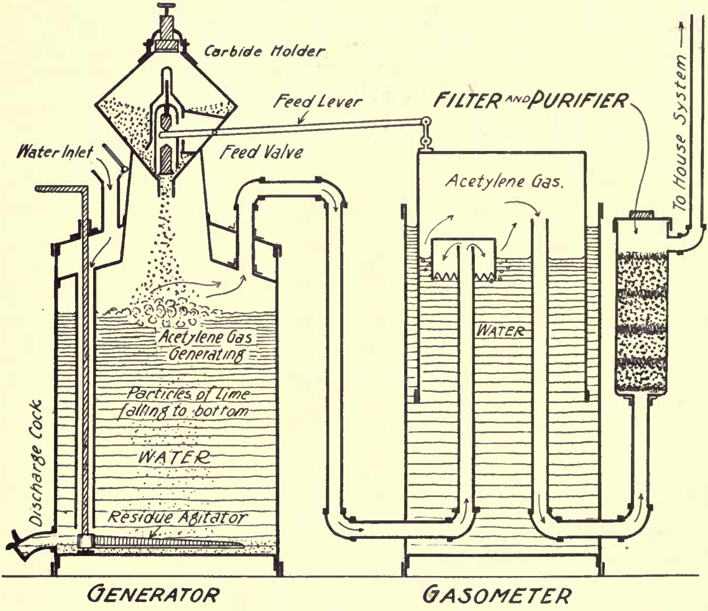

FIG. 71 — Automatic acetylene generator and gas holder.

Alcohols. If one of the hydrogen atoms of methane is replaced with a hydroxyl group, we obtain CH_3OH. Since this is derived from methane, its name should suggest that gas. It is called methyl alcohol. An alcohol is a hydroxide derived from a hydrocarbon. There are hydroxides derived from other hydrocarbons, as ethyl alcohol, C_2H_5OH, and butyl alcohol, C_4H_9OH. Some alcohols contain more than one hydroxyl group, as glycerin, $C_3H_5(OH)_3$.

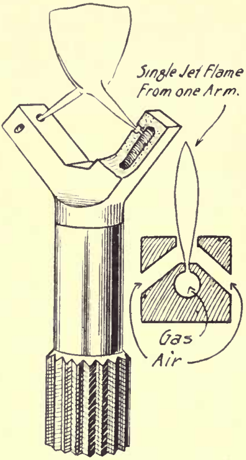

FIG. 72.—Acetylene burner.

Aldehydes. If we gently oxidize an alcohol, we obtain an aldehyde.

$$2\,CH_3OH + O_2 \rightarrow 2\,HCHO + 2\,H_2O$$

This particular aldehyde is called methyl aldehyde. Its trade name is formaldehyde, or formalin. It is used largely as a disinfectant and as a preservative. A small amount is sometimes added to milk to make it keep, but this is prohibited by law, as formaldehyde is a poison.

Organic acids. If the oxidation of an alcohol is carried still further, we obtain an acid.

$$2\,HCHO + O_2 \rightarrow 2\,HCOOH$$

The acid derived in this way from methyl alcohol is formic acid. You are familiar with one such change, the oxidation of ordinary alcohol (ethyl alcohol), to produce acetic acid, or vinegar.

$$C_2H_5OH + O_2 \rightarrow HC_2H_3O_2 + H_2O$$

Esters. In some ways the action of an alcohol on an acid is similar to the action of a metallic hydroxide on an acid.

That is, the organic hydroxide will form a salt with an acid, just as a metallic hydroxide will. These organic salts are called *esters*. If sulphuric acid is mixed with sodium acetate, acetic acid is set free and sodium sulphate is formed. If, now, ethyl hydroxide (ethyl alcohol) is added, ethyl acetate, a liquid having an agreeable fruity odor, results.

$$H_2SO_4 + NaC_2H_3O_2 \rightarrow NaHSO_4 + HC_2H_3O_2$$
$$HC_2H_3O_2 + C_2H_5OH \rightarrow C_2H_5C_2H_3O_2 + H_2O$$

The odor and taste of many flowers and fruits is due largely to the presence of these organic salts. Thus, oil of wintergreen is almost entirely methyl salicylate, mixed with a small per cent of ethyl salicylate. By preparing these two bodies synthetically and mixing them in the correct proportions, an oil is obtained that can hardly be distinguished from the natural oil of wintergreen.

By preparing such esters synthetically, and mixing them in the proper proportions, many flavors and perfumes can be prepared in the laboratory. Usually, however, a natural perfume is a very complex mixture, and its aroma is due to the presence of small quantities of many compounds. To reproduce it perfectly is therefore a difficult task.

Fatty acids. In certain parts of Africa, where sour fruits are not common, the natives regard large white ants as a great delicacy. This is because the ants contain formic acid, which gives an agreeable acid taste to their bodies. This formic acid, $HCHO_2$, is the first member of the series of the organic fatty acids, some of the important members of which are given in the table on page 207.

These acids all contain the group C-O-O-H (carboxyl group), which is characteristic of all organic acids. The first members of the group are thin liquids like water, in which

COMMON ORGANIC ACIDS

NAME:	FROM:	FOUND IN:
Formic acid, $HCHO_2$,	Methane, CH_4,	Ants, stinging nettles.
Acetic acid, $HC_2H_3O_2$,	Ethane, C_2H_6,	Vinegar.
Propionic acid, $HC_3H_5O_2$,	Propane, C_3H_8,	Wood distillate.
Butyric acid, $HC_4H_7O_2$,	Butane, C_4H_{10},	Rancid butter.
Valeric acid, $HC_5H_9O_2$,	Pentane, C_5H_{12}	Valerian root; whale oil.
Caproic acid, $HC_6H_{11}O_2$,	Hexane, C_6H_{14},	Coconut oil.
Caprylic acid, $HC_8H_{15}O_2$,	Octane, C_8H_{18},	Butter and cheese.
Capric acid, $HC_{10}H_{19}O_2$,	Decane, $C_{10}H_{22}$,	Limburger cheese.
Palmitic acid, $HC_{16}H_{31}O_2$,		Butter and tallow.
Stearic acid, $HC_{18}H_{35}O_2$,		Tallow and lard.
Arachidic acid, $HC_{20}H_{39}O_2$,		Peanut oil.
Medullic acid, $HC_{21}H_{41}O_2$,		Ox marrow; beef fat.
Cerotic acid, $HC_{27}H_{53}O_2$,		Beeswax.

they are freely soluble. They show all of the acid properties with which we are already familiar. The higher members are still liquids, but become more and more oily as the number of carbon atoms in the molecule increases, while the highest members are tasteless solids, and show only a weak acidity.

Some other common fatty acids, which do not belong in this series, are: Oleic acid, $HC_{18}H_{33}O_2$; Linoleic acid, $HC_{18}H_{31}O_2$; and Ricissoleic acid, $HC_{18}H_{33}O_3$.

Fats and oils. In nature the higher fatty acids generally occur not free, but combined with glycerin, in the form of esters. These esters are our ordinary oils and fats. The great bulk of oils and fats is composed of the glycerides of oleic, palmitic, and stearic acids. These are called olein, palmitin, and stearin. Olein is a liquid at ordinary temperatures, palmitin and stearin are solids. All three are almost tasteless and odorless.

Olive oil, as an example of an oil, consists mainly of olein and palmitin. Lard, as an example of a fat, has the same composition. Lard, however, is solid because it contains a larger percentage of palmitin than does olive oil. Beef tallow is mainly stearin, and is hard. The consistency of a fat depends upon the proportions of the olein, stearin, and palmitin that it contains.

The differences in the flavors of fats and oils are due to small amounts of other compounds present. Fats and oils cannot be distilled, as, when heated much above their melting point, they decompose, giving off an acrid smoke. The various fats and oils are insoluble in water, but are easily dissolved in such organic solvents as gasoline, ether, and oil of turpentine. Carbon tetrachloride, CCl_4, is largely used as a fat solvent under the trade name of Carbona.

The formulas of the common fats are: Palmitin, $C_3H_5(C_{16}H_{31}O_2)_3$; Olein, $C_3H_5(C_{18}H_{33}O_2)_3$; Stearin, $C_3H_5(C_{18}H_{35}O_2)_3$.

Hydrolysis. When a fat becomes rancid, it is due to the separation of the glycerin and the fatty acid radical, setting free the fatty acid itself. Thus, the glyceryl butrate contained in butter is easily decomposed into glycerin and butyric acid. Butyric acid has a very disagreeable taste and odor, and we say that the butter has become strong. This change is called *hydrolysis,* because water is taken up.

$$\underset{\text{glyceryl butrate}}{C_3H_5(C_4H_7O_2)_3} + \underset{\text{water}}{3\,H_2O} \rightarrow \underset{\text{glycerin}}{C_3H_5(OH)_3} + \underset{\text{butyric acid}}{3\,HC_4H_7O_2}$$

Benzol. *Benzol,* C_6H_6, is obtained from coal tar by distillation. It is the starting point for thousands of synthetic compounds, many of which are used in medicine. Benzol is a colorless, volatile liquid. It burns with a smoky flame, and is one of the illuminants in gas. It is much used as a solvent for sulphur, phosphorus, oils, and rubber. As it readily dissolves grease, it is used to "dry clean" clothes.

Nitrobenzol. *Nitrobenzol,* $C_6H_5NO_2$, is made by the action of nitric acid upon benzol. It is a heavy, oily liquid, having the odor of bitter almonds. It is used as a scent in soap, under the name of oil of mirbane, or artificial oil of bitter almonds.

Aniline. *Aniline,* $C_6H_5NH_2$, is made from nitrobenzol by reducing it with *nascent* hydrogen. Hydrogen, produced by the action of zinc on hydrochloric acid, at the *instant of its liberation,* is very energetic. It is said to be *nascent.* Aniline is a volatile, colorless liquid, which unites directly with acids to form salts, as aniline hydrochlorate, $C_6H_5NH_2HCl$. Toluidin, $C_6H_4CH_3NH_2$, resembles aniline. These two compounds are interesting to us, because their mixture, when oxidized, yields the various colors known as aniline dyes. A slight change in the composition of these dyes changes the color very materially.

Carbolic acid. *Phenol, carbolic acid,* C_6H_5OH, is chemically not an acid at all. It has, however, a corrosive action on flesh, whence its popular name of acid. It is found in coal tar, and is separated from it by distillation. Pure phenol is a colorless, crystalline solid, soluble in 20 parts of water at ordinary temperatures. Its solution is a powerful antiseptic, and is used as a disinfectant. It is poisonous, and should be used with care.

Some important benzol derivatives. *Acetanilide,* sometimes named antifebrine, used as a sedative and for fevers; *phenacetin,* used in headache powders; *naphthalene,* the white crystal sold as "moth balls"; and *creosote,* used as a wood preservative and a medicine, are a few of the benzol series compounds that are of importance to us.

SUMMARY

A hydrocarbon is a compound containing hydrogen and carbon.

An alcohol is an organic hydroxide.

An aldehyde is the product of the partial oxidation of an alcohol, and always contains the group C-O-H.

An organic acid is the product of the oxidation of an alcohol, and contains the carboxyl group C-O-O-H.

An ester is an organic salt.

A fat or oil is a glyceride of oleic, palmitic, or stearic acid.

Benzol is a hydrocarbon obtained from coal tar. Nitrobenzol, aniline, and thousands of other compounds are made from it.

Exercises

1. How would you remove a grease spot from a suit?

2. Rancid butter, when washed with water, becomes edible. Why?

3. How could you make ethyl nitrate?

4. Sodium oleate is soluble in water. If sulphuric acid is added to the solution, a white insoluble solid is set free. What is it? Write the equation.

5. How could you convert an aldehyde into an acid?

CHAPTER XXI

OILS, FATS, AND SOAP

There are so many thousands of organic compounds that no one man knows all their properties and uses. A few of them though are used so largely in our daily life that we should know something of their preparation and uses. Among the most important of these are the oils and fats, used so extensively as foods.

Methods of extracting oils. There are a number of general methods used to extract oils. One that is largely used is pressing, or *expressing,* as it is generally called. The material is placed in bags, and pressed, gently at first, which yields the best oil, and then more forcibly, which yields an oil of a somewhat lower grade. Olive oil is obtained by this process.

Volatile oils, as the oil of cloves, are often extracted by placing the ground material in a still, and then passing live steam into it under pressure. The volatile oil is carried over with the steam, and, condensing with it, sinks to the bottom of the receiver. Minute traces of essential oils may be extracted by this process. Thus it takes 4000 pounds of rose leaves to furnish one pound of the pure oil of rose by steam distillation.

The oil obtained by distillation from flowers does not always have the exact odor of the flower, for the flower odor is a blend of the odor of many substances, not all of which are volatilized with the steam. Then, too, the heat of the steam destroys some of the more delicate odors. To avoid these difficulties, some of the more delicate flower odors are obtained by the *enfleurage* process. A neutral, odorless fat is spread on glass plates, and the flowers strewn on it (Fig. 74). The fat absorbs the odor, and, when the flowers are exhausted, they are replaced by fresh ones. This is repeated

a. Distillation of volatile oils from geraniums.

b. Sorting roses for extraction of oils by the enfleurage process.

FIG. 73. — Preparation of essential oils.

until the fat has become saturated with the perfume of the flower. From this fat the odor is extracted with alcohol.

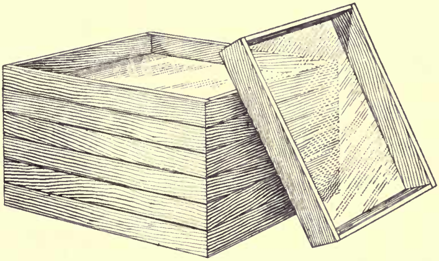

FIG. 74. — Trays used in extraction of volatile oils by the enfleurage process.

Some oils are obtained by the use of solvents. The ground material is placed in closed vessels and treated with an oil solvent, as benzine. The solution obtained is then distilled to recover the solvent. The yield is larger than that obtained by pressing, but the apparatus used is costly, and the press cake obtained cannot be used as a cattle food. The process is therefore used mainly to obtain the essential oils from flowers, where the material to be handled is not large in bulk, and the product obtained is high in price.

Fats and oils classified. There is no chemical difference between fats and oils. If they are liquid at ordinary temperatures, we call them oils; if solid, fats. Even this rule is not of universal application, for palm oil is a solid. Waxes, however, belong to a different class of compounds. They are as a rule fatty acid radicals united with an alcohol of some other than the glycerin series. Chemically, the paraffin oils are not oils at all, but hydrocarbons.

For our purpose, fats and oils may be classified as *drying oils*, such as linseed oil, which, exposed to air, oxidize and become solid; *semi-drying oils*, which partially oxidize; *non-drying oils*, which do not change on exposure to the air, as olive oil; and *volatile oils*, such as oil of cloves, that readily evaporate. The first three classes are sometimes called *fixed oils*, because they are non-volatile.

Linseed oil. Linseed oil is obtained by pressure from the ground seeds of the flax plant. It is light yellow in color. Its drying properties make it valuable in paints and varnish. Linoleum is made by mixing ground cork with linseed oil. The oil dries, that is, oxidizes and becomes a solid, thus binding the particles of cork together. To hasten this oxidation, the oil is heated with salts of manganese or lead. These compounds are known as *driers*. The oxidation of linseed oil, and other drying oils, generates heat, and is often a cause of spontaneous combustion.

Cottonseed oil. Cottonseed oil is one of the semi-drying oils. It is obtained by expression from the cotton seed left from ginning cotton. The press cake left after the oil is pressed out is used as a cattle food. The oil at first is almost black. To purify it, the oil is agitated with a small amount of dilute sodium hydroxide solution. This removes the color and free fatty acids. On standing, the pure yellow oil separates, is drawn off, and washed with water.

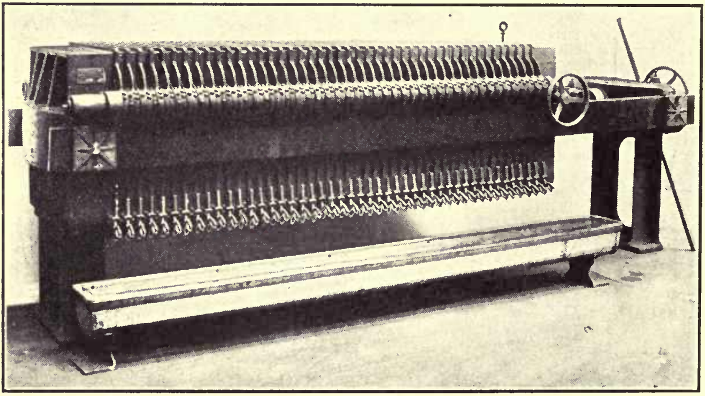

Fig. 75. — A filter press.

Cottonseed oil is used in the manufacture of soap, and for edible purposes. The "salad oil" of the grocer is likely to be cottonseed oil. On cooling the oil, stearin separates as a light yellow solid of the consistency of butter. This is filtered out and used largely in making lard and butter substitutes.

Sesame oil. Sesame oil is obtained from *Sesamum orientale,* grown in India, China, and West Africa. It has a pleasant taste, and is used to some extent in cooking. It is cheaper than olive oil, and can be used instead of it for many purposes in the household. It deserves a more extended use.

Peanut oil. Peanut oil is obtained from the shelled nut of the common peanut by pressure. The cold pressed oil has a pleasant flavor and is used as a salad oil. A second quality obtained by hot pressing is darker in color, and is used in the manufacture of soap. Peanut oil is a typical non-drying oil, that is, it does not become thick because of oxidation on exposure to the air.

Olive oil. Olive oil, used so extensively as an edible oil, is obtained by crushing and pressing ripe olives. The ripe olive is dark in color, rich in oil, and very different from the bottled olives with which you are familiar. The first gentle pressing yields the "virgin oil," which is the best grade of table oil. The pulp is then treated with water and pressed again. This gives a slightly inferior oil, which is used in cooking and for salads. A final treatment with hot water and pressure yields an oil suitable only for soap making. It is from this oil that castile soap is made. Olive oil is a non-drying oil.

Other vegetable fats. Several other vegetable fats are of importance. Palm oil, obtained from the fruit of palm trees, and coconut oil obtained from the coconut, are largely used in making soap. Cocoa butter is the fat pressed from chocolate in the manufacture of cocoa, and is used in toilet preparations, ointments, and in confectionery.

How animal fats are extracted. Animal oils and fats are "rendered." The fat is cut into small pieces and thrown into a kettle, where it is heated with live steam. After the fat has melted, the content of the kettle is filtered to remove the animal membranes, and the liquid allowed to stand. The oil rises and is removed. The watery remainder is used with a second portion of fat. Beef tallow, mutton tallow, and lard are obtained in this way.

Butter and lard substitutes. Butter substitutes, called *butterine* or *oleomargarine*, are made by rendering beef fat at a low temperature. This frees it from the animal membranes. The clear fat is then cooled and kept for some days at a temperature of about 80° F. in order that the stearin contained in it may separate. The fat is then placed in bags, and subjected to a heavy pressure to separate the stearin from the oleo oil. This oleo oil that runs out from the bags is a clear, tasteless, light yellow fluid, which solidifies, on cooling, to a crumbly mass. White stearin is left in the bags, and is used to make candles.

This oleo oil would not be an acceptable substitute for butter, as it would lack the butter flavor. This flavor is imparted to it by churning it with milk, at a temperature that will keep the oleo oil melted. This adds a small amount of butter to the fat, and gives it an agreeable flavor. To make the flavor of the butterine more like that of butter, small amounts of propionic, butyric, and capric acids are sometimes added. Cottonseed oil or cottonseed stearin are also often added.

To give it a grain similar to butter, the mass is then cooled suddenly by running it into ice water. It is then worked to free it from water; it is salted and colored in the same way that ordinary butter is. As in winter the product would be too hard to resemble butter, some soft fat, as cottonseed oil, is added.

Butterine has an advantage over butter in that it is much cheaper, and keeps better, as it contains less of the easily decomposed glycerides. It is probably somewhat less digestible than pure butter, but not enough so as to render its use objectionable. To sell oleo as butter is of course a fraud, but to sell it on its own merits, as a cheap butter substitute, is not objectionable.

Mixtures of beef and cottonseed oil are largely used in cooking under such names as Cottolene. The beef fat in this is expensive, and chemists have lately found a way to add hydrogen to cottonseed oil, thus converting it into a hard solid. This is used in Crisco and other lard substitutes instead of the more expensive beef fat.

Manufacture of oilcloth. The drying qualities of linseed oil are taken advantage of in the manufacture of oilcloth and linoleum. To make oilcloth, burlap is first sized, so as to fill up the pores. It is then heated, to dry it. This prepared cloth is then thickly covered with red lead paint, which is linseed oil mixed with red lead. This painting is repeated five times, the cloth being heated after each painting to hasten the drying of the oil. The surface is then rubbed with pumice, to make it even. The design is printed on the surface, using a special form of printing press. The cloth is again heated, and then taken to the varnishing machine. Here the glossy finish is given it by flowing on it a coat of varnish. A final heating to dry the varnish finishes the operations.

If a thin table or shelf oilcloth is to be made, the process is much the same. Cotton sheeting is used as the foundation, and this is coated with a mixture of linseed oil and china clay. This may be colored to give any desired tint. The design is printed on the finished oilcloth in a cylinder press, similar to those used in calico printing. It is then varnished, heated for 24 hours to dry it, and is then ready for use.

Manufacture of linoleum. Linoleum is made by boiling linseed oil in large kettles, with the addition of driers, until it becomes thick. Canvas sheets are then dipped in the oil, and hung up. The sheets are flooded with oil twice a day, and, being kept in a room the temperature of which is 165° F., the oil rapidly hardens.

In two or three weeks the mass of hardened oil has become so thick that the canvas sheet resembles a thin board. The " skins " are now cut down, and ground up into flakes. These flakes are mixed with powdered cork, and wood pulp, and any color desired is added. The mass is then passed between rollers, which press it into a sheet 18 inches wide by $\frac{1}{8}$ inch thick. These sheets have the consistency of soft putty. Dies now cut these differently colored sheets into blocks. These are laid on a burlap covered with red lead paint, so as to form any desired design. A pressure of 3000 pounds to the square inch is then applied, using a hydraulic press. This consolidates the material into one solid mass. To complete the drying of the oil, the linoleum is then heated to 145° F. for from four to five weeks. Varnishing completes the process, and the linoleum is ready for market.

This is the best grade, called inlaid linoleum. As the colors go through from front to back, they cannot wear off, as they do in oilcloth. A cheaper grade of linoleum is made by printing the design on the surface, instead of inlaying it. Cork carpet is linoleum made of an extra thickness, and only powdered cork is used as a filler. It is expensive, but lasts almost indefinitely.

Saponification explained. When any of the common fats or oils, which are all glycerides of fatty acids, are treated with a metallic hydroxide, a change called *saponification* takes place. The fatty acid radical of the fat combines with the metal, while the glyceryl radical combines with the hydroxyl, producing glycerin.

$$\underset{\substack{\text{glyceryl}\\\text{stearate}}}{C_3H_5(C_{17}H_{35}CO_2)_3} + \underset{\substack{\text{sodium}\\\text{hydroxide}}}{3\ NaOH} \rightarrow \underset{\text{glycerin}}{C_3H_5(OH)_3} + \underset{\substack{\text{sodium}\\\text{stearate}}}{3\ NaC_{17}H_{35}CO_2}$$

If we use sodium or potassium hydroxides, the resulting salt is soluble in water and is called *soap*. By using lead or zinc hydroxides, we can form lead soaps or zinc soaps, but, since these are insoluble, they are of no value as soaps. Zinc oleate is used as the basis for zinc ointment.

Manufacture of soap. Common laundry soap is made largely from animal fats. The tallow or grease used comes to the soap factory in large iron or wood casks. Live steam is used to melt the fat, which is then placed in large iron vats holding several tons. Here sodium hydroxide is added, and the mass heated by steam coils. The boiling is continued until saponification occurs.

A salt solution is then added. As soap is insoluble in brine, the contents of the vat separate into two layers, an upper layer of soap, and a lower layer of brine mixed with glycerin and some impurities. The lower layer is drawn off and the glycerin recovered from it. This is the source of practically all of the glycerin of commerce. The soap remains in the vat. The change is a slow one, the whole process taking about forty-eight hours.

Soda lye is then again added, and live steam introduced until the mass is once more boiling. In laundry soaps rosin is now introduced. This makes the soap lather well, and cheapens the product.

In good grades of toilet soap no rosin is used. About half as much rosin is added as fat was used. Rosin soap is not as good a detergent as a fat soap, but its cost is much less. Sufficient sodium hydroxide must be used to saponify the rosin, as well as to complete the saponification of the fats. When the saponification is complete, brine is again added, and the soap allowed to stand until it has again separated. The lower layer is then drawn off. This requires about twenty-four hours.

A third charge of strong sodium hydroxide is then added, and the mass heated with steam. This is to insure the saponification of all of the fat and rosin. This is called the "strength change," and requires twenty-four hours. The soap is now finished, but it still contains some lye and

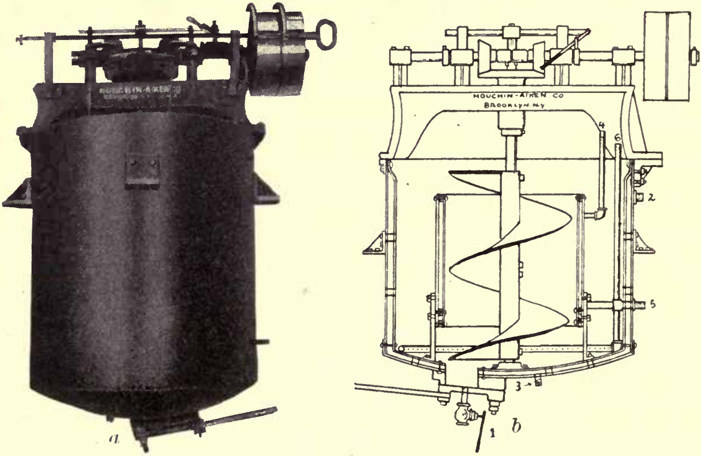

FIG. 76. — A crutcher. *a.* External view. *b.* Section.

impurities. These are removed by boiling with water until the soap loses its granularity, and becomes a smooth mass.

The contents of the vat are now allowed to remain undisturbed for two days, when three layers form. The upper is the pure soap. The next is a dark layer of soap mixed with impurities. This is allowed to remain in the vat until enough has accumulated to make it worth while to bleach it. It is sold as a low grade soap. The bottom layer is lye, which is drawn off and used again.

The soap is run into a mixing machine called a crutcher, where sodium silicate, borax, or sodium carbonate, is added (Fig. 76). Sodium silicate is added to make the soap hard, and prevent its wasting away too quickly. The soap is then run into large wooden frames, where it solidifies. It is cut into cakes, and after these have dried a little, they are stamped with the design of the manufacturer, wrapped, and sold.

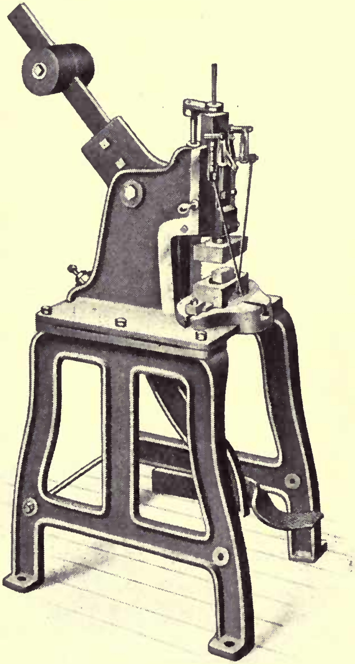

FIG. 77. — Soap stamping machine.

Manufacture of toilet soap. In making toilet soap, the soap, after cooling in the frames, is cut into chips, dried, and run through a second mixing machine called a mill. Here the color and perfume are added. To produce the medicinal soaps, antiseptics, creosote, carbolic acid, and other products are added to the soap in the mill. The soap is then pressed into cakes and stamped (Fig. 77).

Potassium soaps lather more freely than sodium soaps. Shaving soaps are therefore potassium soaps. A little potassium bicarbonate is also often added. In making soft soap, potassium hydroxide is used instead of sodium hydroxide.

Soap fats. The fats used depend on the product desired. The cheaper fats are used in laundry soaps, while the more expensive are used in toilet soaps. Floating soaps are made by beating air into the soap while it is soft. They are no purer than other kinds.

Water in soap. It is to the advantage of the manufacturer to sell soap containing a large quantity of water. A seemingly dry coconut oil soap may contain 60% of water. Soap will last longer, if allowed to stand, unwrapped, in the air after buying it, as this allows some of the water to evaporate. The soap is then harder and does not waste away so quickly. Common laundry soap will lose 25% of its weight.

Soap powders. Soap powders are made by grinding dry soap to a powder, and adding sodium carbonate or borax. If you will weigh the contents of a box of any of the soap powders, and compare its cost with the cost of an equal amount of soap and sodium carbonate, you will find that they are very expensive. You can make an equivalent preparation at home at much less cost.

Scouring soaps. If an abrasive is added to the soap, we have a scouring soap, as the familiar Sapolio. These scouring soaps are usually dried in molds, as they are difficult to cut. Ground quartz, feldspar, or pumice are some of the abrasives added. They also often contain a little sodium carbonate.

Cold process soap. It is not necessary to boil fat to saponify it. In the cold process of making soap, the fat and sodium hydroxide are mixed together and allowed to stand. If alcohol is used instead of water a liquid soap is obtained. Transparent soaps are made by dissolving soap in alcohol and distilling off the alcohol. This leaves the soap as a transparent jelly.

Soaps to be avoided. In buying soap, do not buy a cheap, soft, highly colored, and scented variety. The color and scent were added for the purpose of concealing poor materials. A toilet soap should be neutral, as otherwise it damages the skin.

Action of soap on hard water. When soap is used with hard water, a white precipitate forms. This is lime soap, and represents a considerable waste. To avoid this, the water should be softened before the soap is used. This may be done by boiling, in the case of temporary hard waters:[1]

$$CaH_2(CO_3)_2 + \text{heat} \rightarrow CaCO_3 \downarrow + H_2O + CO_2$$

or, by the addition of washing soda to permanently hard waters:

$$CaSO_4 + Na_2CO_3 \rightarrow CaCO_3 \downarrow + Na_2SO_4$$

SUMMARY

Fats and oils are obtained by expressing, by enfleurage, by dissolving in an oil solvent, and by distillation with steam.

Drying oils, as linseed oil, oxidize on exposure to the air and become solid.

Non-drying oils, as olive oil, do not change on exposure to the air.

Semi-drying oils, as cottonseed oil, thicken on exposure to the air.

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[1] Temporary hard water contains soluble calcium acid carbonate which is decomposed when the water is boiled. Permanent hard water contains calcium sulphate which is not affected by boiling.

Fixed oils, as corn oil, are non-volatile.

Volatile oils, as oil of cloves, readily evaporate.

Oleomargarine, or **butterine**, is made from beef fat mixed with cottonseed oil. A small amount of real butter is mixed with it to give it a butter flavor. It is a cheap, desirable, butter substitute.

Linoleum and **oilcloth** are cloths covered with a very thick layer of oil or paint.

Soap is a sodium or potassium salt of a fatty acid.

Soft soaps are potassium soaps.

Hard soaps are sodium soaps.

Liquid soap is a solution of soap in weak alcohol.

Scouring soaps contain some abrasive, as powdered feldspar.

Exercises

1. How would you tell to which class of oils peanut oil belongs?

2. Would you be willing to use butterine at home? Explain.

3. How could you prepare oil of cinnamon at home?

4. How could you clean a greasy waist?

5. Could linseed oil be used to adulterate olive oil? Explain.

6. Why is oil of rose so expensive?

7. How could rancid butter be made fit for use? Is this renovated butter worth as much as fresh butter?

8. How could you distinguish betweeen a soft and a hard water?

CHAPTER XXII

CARBOHYDRATES

Carbohydrates defined. Nature is the great master chemist of the world, and one of her most marvelous doings is seen in the ease with which she builds up complex organic compounds from the simplest materials, as in the production of starch by green plants. The chlorophyll cells of plants having green leaves are able, under the influence of sunlight, to cause carbon dioxide and water to combine, forming an organic compound and setting free oxygen. Just how the plant does this, we do not yet know. The equation below represents only the end products; of the intermediate steps we are not sure.

$$6\,CO_2 + 5\,H_2O \rightarrow C_6H_{10}O_5 + 6\,O_2$$

You will notice that in the formula of the compound formed, called starch, hydrogen and oxygen occur in the proportion in which they are present in water, and that there are six atoms of carbon. Compounds of which this is true are called *carbohydrates.* In them you will always find oxygen and hydrogen present in the proportion of 1 : 2, and carbon present as six atoms, or a multiple of six. These carbohydrates are very important compounds for our study, as they include the sugars, starch, and cellulose. These compounds are of especial interest in the home.

Cane sugar. Cane sugar, or sucrose, $C_{12}H_{22}O_{11}$, is found in many plants, but in most of them it occurs in too small quantities to pay for extracting. All that is of importance commercially comes from sugar cane, sugar beet, and sugar maple.

Sugar cane is a large, jointed grass, somewhat like corn, numerous varieties of which are grown in tropical and subtropical regions of the earth, as Cuba, Louisiana, and Texas (Fig. 78). The stalk of sugar cane contains from 12% to 20% of cane sugar. When ripe, the cane is cut, the leaves and green top removed, and the juice expressed by passing the cane between rollers, arranged somewhat like those of your clothes wringer. Lime is then added to coagulate the impurities, and to neutralize any acids that may be present. The sweet juice is then boiled, to evaporate some of the water. The scum that rises is removed. When enough water has evaporated, so that the solution is ready to crystallize, it is run into tanks, where it cools, and crystals of cane sugar

FIG. 78. — Growing sugar cane.

are obtained. The liquid that does not crystallize is molasses, which is drained off. The sugar crystals remaining are "raw sugar." Generally the raw sugar is not refined where it is made, but is shipped to a large sugar refinery.

Refining sugar. In the sugar refineries the raw sugar is dissolved in hot water, and filtered to remove any insoluble material. It then forms a clear sirup, but is dark in color. To remove this brown coloring matter, the sirup is passed through large cylinders filled with bone black.

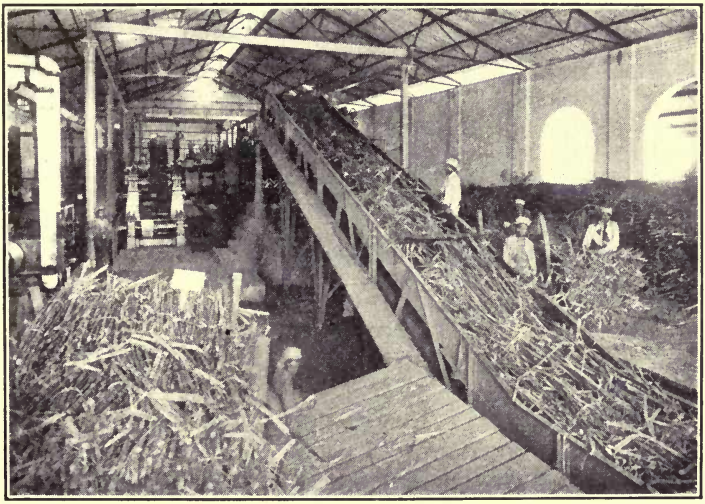


FIG. 79. — View in a sugar cane mill.

When the sirup leaves the bone black filters, it is not only clear, but colorless. It must now be evaporated to crystallize the sugar. Sugar sirup upon boiling undergoes a chemical change called *hydrolysis,* in which one molecule of sucrose combines with one molecule of water and gives one molecule of glucose and one of fructose. Technically this change is called *inversion,* and the product *invert* sugar. Both of these substances are sugars, but they are not as sweet as sucrose, nor do they crystallize as well. It is therefore necessary for the sugar refiner to avoid their formation. To do this, he must avoid heating the sugar to a high temperature. He must, however, boil the sirup in order to evaporate the water and crystallize the sugar.

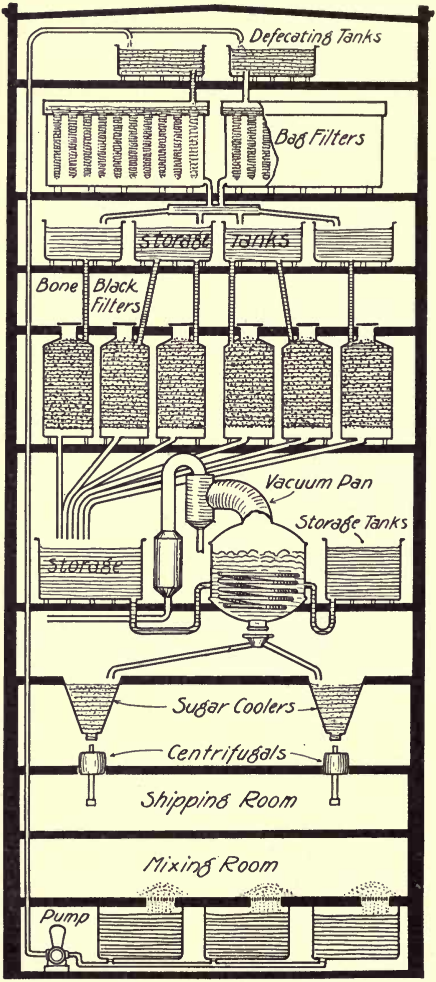


Fig. 80.— Diagram of a sugar refining plant. (After Sadtler.)

It would seem as if this were a case where the manufacturer was bound to find trouble, whichever way he turned, but the solution of the problem is easy. The sirup is run into a large pan, so arranged that it can be closed air tight and then the air is pumped out. We know that water boils at 212° F. at the ordinary atmospheric pressure. If the air pressure is lowered, the boiling point is also lowered. By removing the air from over the sirup, the boiling point of the sirup is lowered, and the water can be evaporated without *hydrolizing* or *inverting* the sugar. It is necessary to pump off the steam as fast as it forms, so as to keep the pressure low. You must remember when cooking fruits that sugar is inverted by boiling, and thus much of the sweetness is lost.

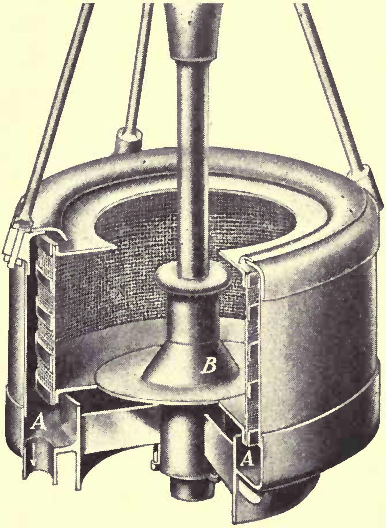


Fig. 81. — Centrifugal machine.

When most of the water has evaporated, the sugar crystallizes, and these sugar crystals must be separated from the sirup. This is done in large *centrifugals* (Fig. 81). You know that when mud accumulates on the rotating wheel of a wagon, there is a constant tendency for it to be thrown off, and that the faster the wheel is rotating, the greater is this tendency. This is due to *centrifugal force,* and this principle is made use of in centrifugal separators, used not only in the sugar industry, but in laundries, creameries, and in many other places.

Imagine two large, tall saucepans, the one fitting inside the other, and the inner one made of wire gauze and so arranged that it can be rotated while the outer one remains stationary. The sugar sirup containing the crystals is put into the inner saucepan, and this is then rapidly rotated.

There is a great tendency for everything in the inner vessel to fly out through the gauze. The sirup can escape through the holes in the gauze, but the crystals are too large to pass through, and so remain in the inner cylinder.

The sugar crystals are then washed by throwing a little water on them, so as to free them from the sticky sirup. The sugar is now pure though it still has a slight yellow tinge. To counteract this, the last wash water is often colored with a little ultramarine. This counteracts the slight yellow tinge, and leaves the sugar a bluish white. You must have seen clothes treated at home in the same way, and for the same reason.

The sugar crystals are now run into the upper end of a large, slightly inclined cylinder, which is heated. As this cylinder is rotated, the crystals tumble over each other until, when they have reached the lower end, they are dry. The granulated sugar is now ready for barreling.

If the moist crystals are pressed together, they form a solid mass that, when broken up, forms lump sugar. If the crystals are ground to a fine powder, we have powdered sugar. If the grinding is continued until a very fine powder results, we call it confectioners' sugar. Rock candy consists of large crystals of cane sugar. It is obtained by allowing sirup to crystallize slowly.

Beet sugar. Sugar cane is not our only source of sugar. Much of the sugar that you use comes from the sugar beet (Fig. 82). In many countries the sugar beet has been cultivated and selected until the best specimens may contain as much as 16% of sugar. The beets are washed, sliced, and put into large iron vats, called *diffusors*. Here water, heated to 140° F., is added, and the sugar diffuses from the beets into the water. The sirup thus obtained is then purified and crystallized in a way similar to that used with cane sugar. More than one half of the sugar produced in the world is beet sugar.

Maple sugar. Many farms in the Northern states and Canada contain a grove of sugar maples, *Acer saccharinum.* In the spring, when the sap starts to flow, the farmer " taps " these trees. That is, he bores a hole through the bark into the wood, and drives a spout into the tree, so arranged that the sap as it oozes out is caught in a pail below. This sap contains about 2% of cane sugar. The sap is then concentrated by heating in large shallow pans until it contains 60% of sugar, when it is poured off to be used as maple sirup. Or, it is concentrated until, on cooling, it solidifies, which gives maple sugar.

Fig. 82. — Sugar beets. (*a*) Wild. (*b*) Cultivated.

Maple sugar is essentially cane sugar, containing a little invert sugar, and certain substances that give it its characteristic taste and color. It is easily adulterated by using cane sugar and a flavoring extract obtained from hickory bark. About 6000 tons are produced annually in the United States.

Fig. 83. — Collecting sap for making maple sirup and sugar.

Sorghum sugar. Small quantities of sucrose are also made from the sorghum cane. In spite, however, of extensive experiments by the government to improve both the cane and methods of extraction, the production of sugar from this source has never been successful commercially.

Boiling sugar. Most of us have a sweet tooth, and a paragraph that will help us to enjoy some toothsome dainties may be welcome.

When sugar is dissolved in a small amount of water, and the solution boiled, the sugar tends to grain or crystallize, as the water evaporates. When we wish to prevent this, we must be careful not to stir or jar the solution. The sirup in contact with the sides of the pan becomes slightly more concentrated than the rest, and therefore crystallization starts there. If these crystals are allowed to remain, they will speedily cause the entire mass to become crystalline; they must therefore be removed by wiping them off with a damp cloth. The addition of half a mustard spoonful of cream of tartar, or of a few drops of lemon juice, to a pound of sugar will help to prevent this crystallization.

The temperature of the boiling sugar solution slowly rises as the water evaporates. When the temperature reaches 215°–217° F., a little of the sirup, when pressed between the thumb and finger, can be drawn out into a thin thread. At this stage the sugar is suitable for boiled icings. At 236°–238° F., a little, dropped into cold water, forms a soft ball that can be rolled between the fingers. This stage is used for fondant. At 310° F., dropped into water, it forms a mass that breaks easily, and is crisp. At this point, called the crack, it crystallizes very easily, so that it is well to add four drops of lemon juice for each pound of sugar. This stage is used for glacé nuts. At 345°–350° F., it turns yellow-brown, and acquires the flavor of caramel. It now contains practically no water, and burns very easily.

These four stages — the thread, the soft ball, the crack, and caramel — are the four most used in cooking. The experienced cook can easily recognize them by the behavior of the sirup, but the use of the thermometer will enable even the beginner to recognize them with certainty.

Maltose. When the ptyalin of the saliva acts upon starch, *maltose* is formed. This sugar has the same formula as sucrose, with the addition of water, $C_{12}H_{22}O_{11} \cdot H_2O$. Maltose is made from starch by the action of malt extract. When seeds germinate, the ferment *diastase* that they contain changes their starch into maltose, on which the young plant feeds. When maltose is inverted, only glucose results.

Lactose. Lactose, milk sugar, $C_{12}H_{22}O_{11} \cdot H_2O$, is the substance that gives milk its sweet taste. Cows' milk contains about five per cent of lactose. In making cheese, the whey of the milk, which contains the milk sugar, is separated. This whey is treated with chalk and aluminium hydroxide, is then filtered, and the clear filtrate evaporated. On standing, the milk sugar separates.

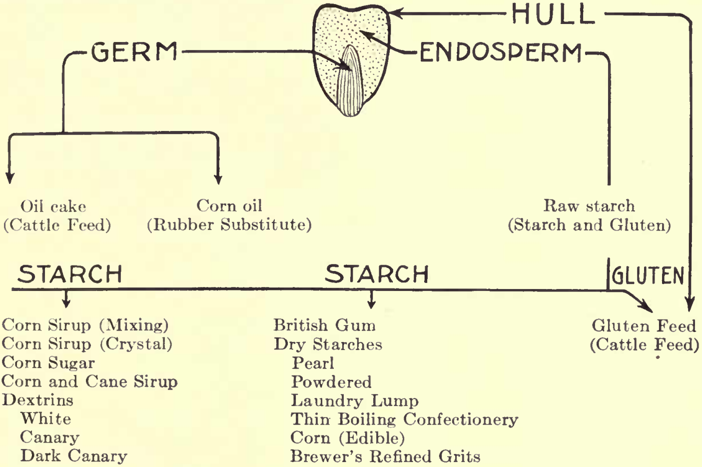
The souring of milk is due to the action of bacteria which change the lactose to lactic acid.

Milk sugar is not as sweet as cane sugar, and is much less soluble in water. It is used extensively in prepared infants' foods, and in the pills and powders of the druggist.

Corn products. Under the general name of *glucose*, a number of different carbohydrates, mixed in varying proportions, are on the market. They are all made from starch by hydrolysis. More than 50,000,000 bushels of corn are annually used in the manufacture of starch and the glucose products made from it. The process is interesting, as it well illustrates how in a modern industry chemistry has enabled us to utilize all of the products obtained. It is this utilization of what was formerly thrown away, that marks one great result of modern chemistry. The following diagram shows the products made from the corn kernel.

The corn is first soaked in water. This softens the grain, and dissolves some soluble materials. This water is not thrown away, but is evaporated, and the residue used in making a cattle food.

The softened corn is then ground coarsely, so as not to crush the germ, and then thrown into water. The germ, being light, because of the oil that it contains, floats, and is taken off. The germs are then dried and pressed, and corn oil is obtained. This is used in making soaps and in oilcloth, and, when purified, in cooking or as a salad oil. The cake remaining in the press is ground and used as a cattle food.



The hulls are then separated from the endosperm, that white part of the grain containing the starch, dried, and ground. A cattle food is made from these hulls, mixed with the gluten which is produced later, and the residue left on evaporating the softening water.

The endosperm is finely ground, mixed with water, and made to flow through long, shallow, slightly inclined troughs.

Here the starch and gluten separate, because of their different specific gravities. The gluten is used as a cattle food.

The starch is dried, and sold as edible corn starch, and as the various grades of laundry starch. The pearl, crystal, and lump laundry starches are all made from the same stock, their different appearances being caused by different methods of drying.

Manufacture of dextrin. If this starch is roasted, dextrin is produced. The different varieties of dextrin, as white, yellow, and British gum, are due to heating to a higher or lower temperature, and by varying the time of treatment. The dextrins are used as adhesives, as in the gum on the back of postage stamps, as sizes in the textile industry, and for thickening the colors used in calico printing. The crust of bread is sweet because of the dextrin that it contains.

Manufacture of glucose. To prepare glucose, starch is mixed with water, a small amount of hydrochloric acid added, and the liquid heated under pressure. The starch hydrolizes, and is converted into dextrin and glucose.

$$(C_6H_{10}O_5)n + n\ H_2O \rightarrow (C_6H_{12}O_6)n$$

If a table sirup is to be made, the conversion is stopped when a product containing 50% glucose and 50% dextrin is obtained.

The hydrochloric acid is next neutralized by the addition of sodium carbonate, the liquid filtered, decolorized by passing it through bone black filters, and concentrated to the desired thickness.

If a solid glucose sugar is to be made, the inversion is continued until most of the dextrin is converted into glucose. It is then purified as for sirup, and evaporated until it solidifies on cooling.

Glucose is used extensively in confectionery, in jams and jellies, pastes and sizes, tanning of leather, and in making vinegar. It is only three fifths as sweet as cane sugar, but has the advantage over cane sugar of being absorbed directly into the body, requiring no digestion. It is a desirable food, but should be sold under its own name, and not under the many fanciful names with which it appears in commerce.

Glucose, dextrose, and grape sugar occur in nature in many fruits. The hard white particles found in raisins are glucose. Fructose, also called levulose and fruit sugar, has the same formula as glucose, $C_6H_{12}O_6$, but its physical properties are somewhat different. It is found in honey.

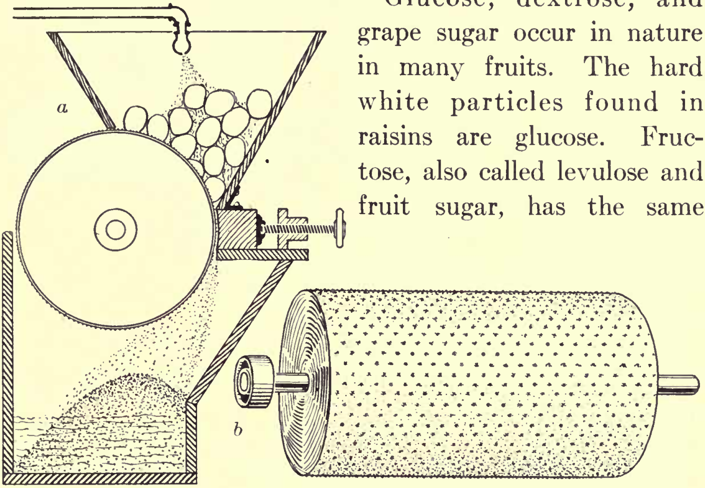


FIG. 84. — Manufacture of potato starch. (*a*) Rasping machine, (*b*) Rasping cylinder.

Starch. Just as business men store up a certain amount of money against a time of need, so plants store up a reserve of food. This accumulation of food is largely in the form of starch ($C_6H_{10}O_5$). It is found in the roots, bulbs, tubers, and sometimes in the stems and leaves.

It is easily separated from the potato by grating the tuber, placing the pulp in a cloth bag, and washing in water. The fine grains of the starch pass through the cloth and make the water milky, while the fiber is held back. On settling, a layer of starch is obtained.

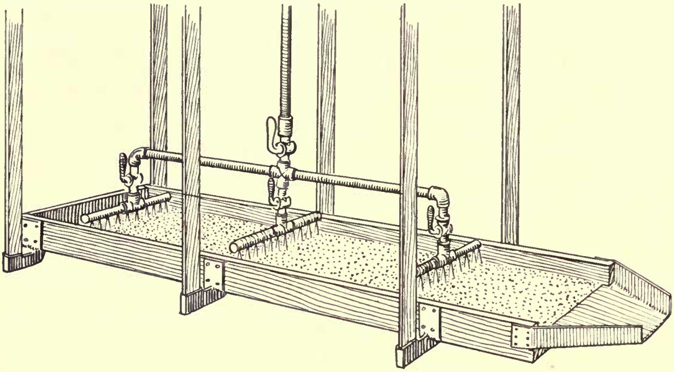


Fig. 85. — Shaking table for separating starch from potato pulp.

Starch grows in fine grains, the size depending on the plant from which it is obtained (Fig. 86). The potato starch granule has an average diameter of $\frac{1}{300}$ of an inch, the granules of wheat starch have an average diameter of $\frac{1}{1000}$ of an inch, and the granules of rice starch are still smaller. The

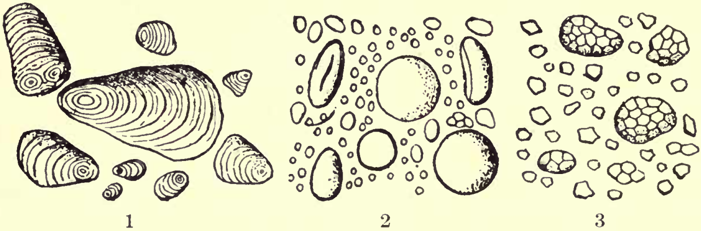
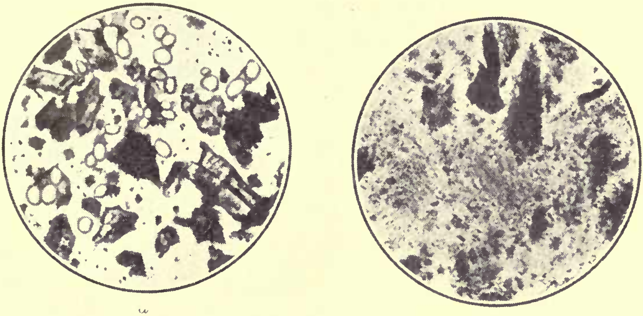


Fig. 86. — Different kinds of starch grains. 1, Potato. 2, Wheat. 3, Rice.

identification of starch under the microscope makes it possible to detect many forms of food adulteration.



a

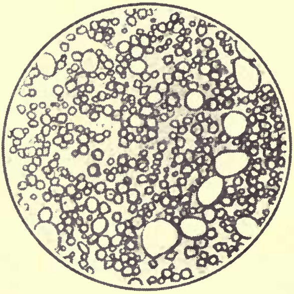
FIG. 87. — (*a*) Pepper starch (angular bodies) adulterated with bean starch (rounded).

b

(*b*) Characteristic sap vessels in chicory by which its presence in coffee is detected.

Starch is not soluble in water. On heating the grains with water to 158° F., the granules swell, and form a kind of semi-solution, as in the case of laundry starch or starch paste.

When heated with an acid, starch hydrolizes, as we have already seen in the manufacture of glucose. This change of starch to sugar also takes place in seeds when they germinate. The change is caused by a ferment contained in the seed, called *diastase*. Every woman is familiar with the uses of starch in making foods, in baking powder, where it is used as a filler, for laundry purposes, and as a size for stiffening fabrics in the textile industry.



c

(*c*) Potato starch (rounded bodies) adulterated with corn starch (angular bodies).

Courtesy of the *Scientific American*.

The saliva contains a ferment called *ptyalin.* When starch is brought into contact with this, it is hydrolized, with the formation of dextrin and maltose. Hence the importance of thoroughly masticating all starchy foods.

Commercial forms of starch. Starch appears on the market not only under its own name, but as sago, tapioca, arrowroot, etc. These are all starches obtained from some particular plant. Sago is made from a palm. In manufacture, it has been heated so that it is difficult to find separate starch grains in it.

Tapioca is made from the cassava plant. The starch is first separated from the woody fiber of the plant by grinding and washing in water. Before the starch becomes dry, it is heated, with the result that the starch grains are broken up, and the starch is left in the familiar form in which we find it in the market.

Arrowroot is usually obtained from a plant called *Canna indica,* which grows mainly in tropical regions. It is also obtained in Bermuda from the *Maranta arundinacea.* These arrowroots are used largely in invalid foods. They form a firm, semi-translucent mass when heated in a small quantity of water. The name comes from the fact that the bruised stems of the plant were used as a poultice for wounds caused by arrows.

Cellulose. Cellulose forms the chief part of wood. It is the building material of all plants, and comprises the cell wall. Absorbent cotton and the best grades of filter paper are nearly pure cellulose.

Pure cellulose is white, insoluble in water, and permanent in the air. Treated with a mixture of strong nitric and sulphuric acids, cellulose is converted into a nitrate of cellulose. This is guncotton, a violent explosive used in torpedoes.

One of these cellulose nitrates is soluble in a mixture of alcohol and ether, and forms collodion. This is used in photography and surgery. If a small amount of castor oil is added to collodion, it remains flexible when it dries, and is then used to cover cuts in the skin.

Celluloid. By adding camphor to guncotton, an elastic mass called celluloid is obtained. When hot, this is plastic, and can be rolled out into sheets, or formed into combs, knife handles, and a multitude of materials used in the home. It can be colored, and made either transparent or opaque. In using celluloid, do not forget that it contains guncotton, and will therefore burn with great violence. Moreover, since it is a nitrate and contains a large amount of oxygen, the flame once started requires no outside supply of oxygen. Burning celluloid therefore cannot be extinguished by smothering. Moving picture films are made on thin sheets of celluloid, and the many disastrous fires that have occurred in moving picture theaters have been due to the difficulty with which burning celluloid is extinguished.

Manufacture of handmade paper. The use of some form of paper dates from the earliest antiquity. We have papyrus rolls, written on by the Egyptians, that date from 4000 B.C. This kind of paper was made from the papyrus reed that grows on the banks of the Nile. The Chinese have long used a paper made from the inner bark of the mulberry tree.

In the Middle Ages, paper was made from linen by mixing the linen fibers with enough water to form a thin, milklike liquid. This was poured on a wire sieve of the size that they wished the finished sheet to be. The water ran through, but the fibers remained in the sieve. This sheet was then placed on a piece of felt and pressed. When dry, a sheet of paper resulted.

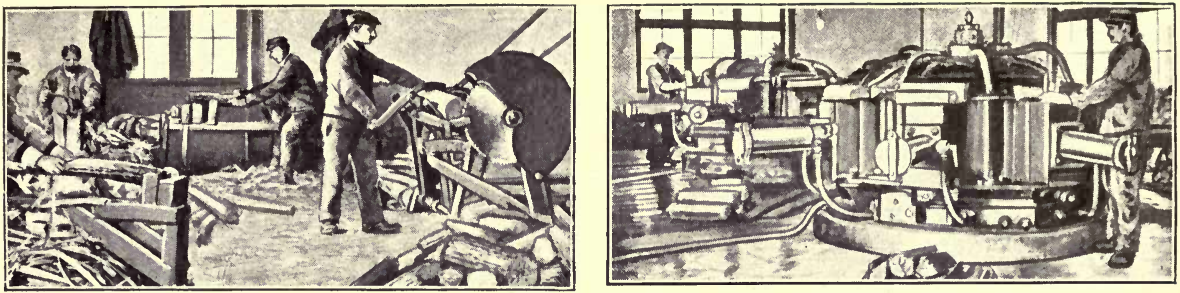
Paper made in this way was porous, like blotting paper, and could not be used to write on, as the ink would spread. To overcome this, the dried sheet was passed through a thin solution of gelatin, to size it. This process is slow and expensive, and to-day only the most expensive papers are made in this way.

Manufacture of machine paper. Modern paper is made on the Fourdrinier machine. The crude material used is rags, wood, esparto grass, old paper, cotton refuse, and many other fibrous materials. The raw stock used depends on the grade of paper wished. The all-rag papers are the best, but also the most expensive. Newspaper is made from wood, and is the cheapest paper made.

In making a rag paper, the rags are first sorted by hand, the buttons cut off, and the cloth cut into small pieces. It then goes to the duster, where it is beaten to remove as much dirt as possible. Here it may lose as much as five per cent of its weight. The rags then pass to the rag boiler, where they are treated with sodium hydroxide or with lime. Much of the coloring matter is here removed while the rags are washed. They then go to the beating engine, where knives cut them up and the stirring separates them into individual fibers. Here also the pulp is bleached, or colored, as may be required.

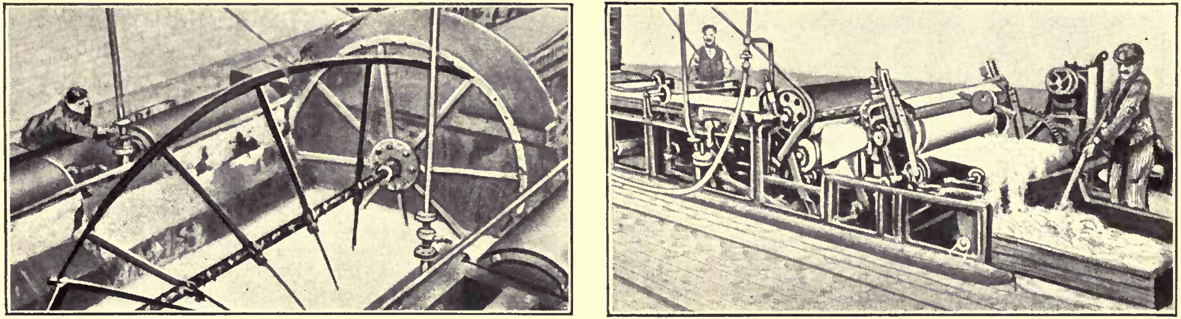
This pulp, mixed with much water, is flowed upon an endless woven screen of wire, which travels rapidly under the box containing the pulp. The water drains through, while the fibers felt together to make an endless sheet of paper. Seventy feet of fine paper, or 600 feet of newspaper, are made a minute.

As the paper passes on, it passes between rollers that squeeze out the water. At the turning point of the endless



Removing bark and sawing logs.

Grinding machines.



Receiving vat.

The pulp press.

FIG. 88. — Manufacture of wood pulp.

(Courtesy of *Scientific American*.)

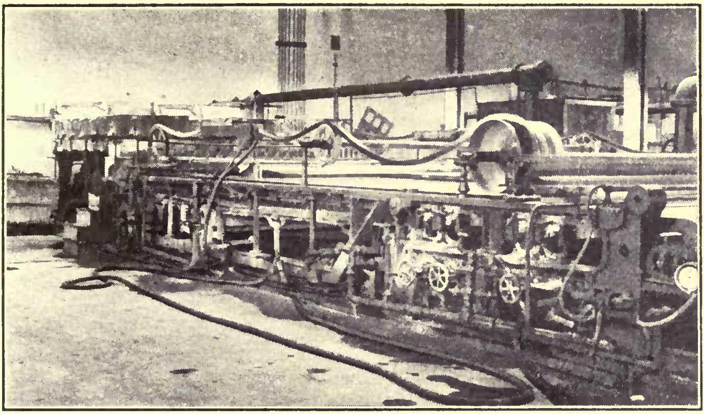
wire belt, the paper leaves it, and passes over a number of hollow iron cylinders, heated by steam. Here it is dried. It then passes between heavy polished steel rollers to smooth it, after which it is cut into sheets. This gives a porous paper.

If a writing paper, or a paper on which to print illustrations containing fine lines is needed, the paper must be sized. Glue or rosin, and some filler, as clay, is then added to the paper pulp. When such a sized paper passes through the last, or calendering rolls, the paper is given a smooth or even a glossy surface. So much clay is sometimes added that, if such " supercalendered " paper is bent, it will crack. You have probably noticed the disagreeable odor that some paper gives off when damp. This is because of the glue used in the sizing.

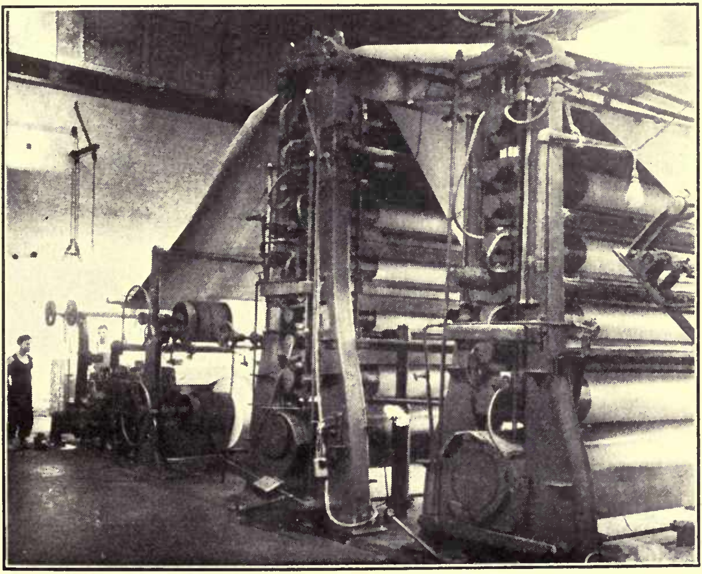
Wood pulp paper. Rag paper is too expensive to be used in newspapers. For this purpose paper made from wood pulp is used. Wood pulp is of two kinds, mechanical and chemical.

Mechanical pulp is simply ground wood. This is made by splitting poplar or spruce into small pieces and grinding them against a wet grindstone. The wood forms a powder fine as flour. The fiber is so short that it would not make paper by itself, but it is useful as a filler.

Chemical wood pulp is made by chipping wood, and then heating these chips in solutions of sodium hydroxide or lime sulphite in large vats under pressure. These chemicals dissolve the materials that hold the wood fibers together. After the " cooking " is complete, which takes from six to eighteen hours, the mass is thrown into water. It is then screened, to remove lumps, and used in the same way that rag pulp is.



a. Paper making machine.



b. Drying and finishing the paper.

Fig. 89. — Paper making.

Common newspaper contains about 25% of chemical and 75% of mechanical wood pulp. Sulphite wood pulp is largely used in book papers. Blotting paper contains no size. Filter paper is made from the best grade of linen fiber, and contains no size. The finish put on some writing papers, as linen finish, is made by passing the paper between rolls that have the desired design engraved on them.

Saccharin. Saccharin, $C_7H_5O_3SN$, is not a sugar, but is a compound that has an exceedingly sweet taste. It is about 500 times as sweet as cane sugar. It is sometimes mixed with glucose in cheap candy to make the product as sweet as if cane sugar had been used. Such a use is of course fraudulent. It is also used by diabetic persons who must avoid the use of sugar. It has no nutritive value and is considered a poisonous and deleterious ingredient in foods.

Carbohydrates as foods. The carbohydrates are very largely used as foods. In nature, they occur in both soluble and insoluble forms. The soluble forms, as the sugars of fruits and honey, require little action of the digestive juices; indeed, many of these soluble forms require none, but are absorbed unchanged by the body to be used in giving heat and energy. The insoluble forms are mainly starch and cellulose. Before these can be made available in our bodies, they must be made soluble. Our digestive juices are not able to do this to cellulose, and so this is unavailable as a food for us. In our digestive tract starch can, however, be made soluble, and it forms one of our most used foods. Cooking starchy foods starts this action by breaking up the starch grains, and, if prolonged, converts some of the starch into other soluble carbohydrates, as dextrin and glucose, thus aiding the action of the digestive juices.

Proteins. One large group of compounds found in all animal and most vegetable matter is the *proteins*. They contain carbon, oxygen, hydrogen, nitrogen, sulphur, and sometimes phosphorus. Nitrogen is the essential element which distinguishes the proteins as a class of food compounds. Albumin, casein, gluten, gelatin, and peptone are all examples of proteins.

Ptomaines. When certain bacteria act upon proteins, they cause putrefaction, forming *ptomaines*. These ptomaines are violently poisonous bodies, and the eating of meat containing them causes ptomaine poisoning, which often results fatally. The heat of summer accelerates the growth of bacteria, and hence causes the quick decay of food. This is the reason why ptomaine poisoning is so much more common in summer.

Albumin. *Albumin* occurs in many of the soft parts of the bodies of animals. The white of egg is almost pure albumin and water. The blood also contains it. It is soluble in water, and is coagulated by both heat and alcohol.

Casein. *Casein* is found in milk. It resembles albumin, but is not coagulated by heat. It is prepared by passing sweet milk through a centrifugal cream separator, to free it from butter fat. An acid is then added, which precipitates the casein in the form of a curd. It is then washed, to free it from milk sugar and acid, and dried. It keeps indefinitely, if kept dry.

It is used extensively as an adhesive, in buttons, in paper making, in casein paints, and in preparing plastic masses. Paper bottles, used to contain milk, are made by soaking paper in a solution of casein, and then exposing it to the vapor of formaldehyde. This makes the casein waterproof.

Casein paints are made by dissolving casein in borax or sodium hydroxide, and adding a filler and a pigment. Clay, lime, or powdered feldspar are used as fillers, and lamp black, cobalt blue, or chrome green as pigments. It is necessary to pick out a filler and a pigment that will not be affected by the alkali used. These paints dry hard, may be rendered waterproof with formaldehyde, and are cheap.

Gelatin. *Gelatin* is an animal jelly. It is the chief constituent of glue. It is made by heating the bones and skins of animals in water for a long time. These animal products contain collagen and ossein, which, on long boiling, change to gelatin. The solution obtained is cooled, when it gelatinizes. The water is driven off, and dry gelatin remains.

Gluten. If you will place a handful of flour in a cloth bag, and knead it under water, the starch grains will pass through the cloth and make the water milky. There will remain in the cloth a tough, yellowish substance called gluten. Some flours, as those used to make macaroni, contain a large percentage of gluten. It is gluten that makes it possible to make bread light. When flour is mixed with water, it is the gluten that forms the pasty dough that entangles the bubbles of carbon dioxide gas.

Alkaloids. Most of the nitrogenous compounds found in nature are foods. One exception is the class of bodies known as *alkaloids*. These are feeble bases, and combine with acids to form salts, which as a rule crystallize easily. They are slightly soluble in water, and readily soluble in alcohol.

The alkaloid, *theine* or *caffeine*, present in tea and coffee, is the stimulating substance found in these beverages. *Theobromine*, a similar alkaloid, is found in the cocoa bean. *Nicotine*, found in tobacco, is intensely poisonous. *Cocaine* is prepared from the leaves of the coca plant. It is used by physicians in the form of *cocaine hydrochlorate* as a local anæsthetic; that is, if applied to some part of the body, as the gums, it makes that part of the body insensible to pain. *Quinine* and *cinchonine* are obtained from the bark of the cinchona tree, and are used as specifics in malaria. All the above more or less poisonous bodies are alkaloids.

SUMMARY

Carbohydrates are compounds containing six atoms of carbon, or a multiple of six, and hydrogen and oxygen in the proportion in which they occur in water. Sugar, starch, and cellulose are important carbohydrates.

Sucrose comes from the sugar beet, the sugar cane, and from the sugar maple.

Glucose, or **grape sugar**, is made from starch by heating it with a minute quantity of acid.

Fructose, or **fruit sugar**, is found in honey.

Maltose is malt sugar.

Lactose is milk sugar.

Hydrolysis is the splitting up of a molecule caused by its combining with a molecule of water.

Inversion is the hydrolysis of a sugar solution. *Invert* sugar is produced.

Dextrin is made by heating dry starch. It is used as an adhesive.

Paper is made from cellulose.

Saccharin is not a sugar. It is 500 times as sweet as cane sugar. It has no nutritive value and is harmful.

Cellulose forms part of the woody part of plants. Absorbent cotton is nearly pure cellulose.

Proteins are nitrogenous organic bodies. Albumin and casein are examples.

Exercises

1. Why does the crust of bread taste sweet?

2. Is there any difference between caramel and burnt sugar?

3. In cooking cranberries, when should the sugar be added and why?

CHAPTER XXIII

FOODS

Food defined. One great difference between animals and plants is in the food that each can use. Plants are able to build up from carbon dioxide, water, and mineral salts, the most complex compounds. Animals cannot do this, but are dependent for food on products that plants have elaborated for them. Plants can get along without animals, but animals cannot get along without plants.

The requirements of our bodies are similar to those of other animals. We need food to repair the cells, to build up new tissue, and to furnish energy and heat to the body. Food may, therefore, be defined as anything which will build up body tissue, and furnish it with heat and energy. This food must be both organic and inorganic.

Inorganic foods. As about 65% of the weight of the body is water, we require large amounts of this to replace the loss that is continually taking place. We take large amounts of water, both directly as a beverage, and in the food that we consume.

The skeleton of the body is largely calcium phosphate, and the teeth, blood, and tissues all contain compounds of iron, sodium, and chlorine. The total of the mineral constituents is about 6% of the weight of the body. This mineral matter is supplied to us in our food. Almost any diet will give more than we require, for vegetables all contain mineral matter that they have taken from the soil.

MINERAL CONSTITUENTS OF MEATS, FRUITS, AND VEGETABLES IN PERCENTAGE OF THE EDIBLE PORTION[1]

Food	CaO	MgO	K_2O	Na_2O	P_2O_5	Cl	S	Fe
Apples	.014	.014	.15	.02	.03	.004	.005	.0003
Beets	.03	.033	.45	.10	.09	.04	.015	.0006
Cabbage	.068	.026	.45	.05	.09	.03	.07	.0011
Carrots	.077	.034	.35	.13	.10	.036	.022	.0008
Cauliflower	.17	.02	.27	.10	.14	.05	.085	——
Grapes	.024	.014	.25	.03	.12	.01	.024	.0013
Lentils	.15	.17	1.05	.08	1.00	.05	.28	.0086
Meats	.011	.03	.3	.057	.33	.03	.13	.002
Muskmelons	.024	.020	.283	.082	.035	.041	.014	0003
Parsnips	.09	.07	.70	.01	.19	.03	.057	——
Spinach	.09	.08	.94	.20	.13	.02	.041	.0032
Squash	.02	.01	.17	.002	.005	.02	.014	.0008
Strawberries	.05	.03	.18	.07	.064	.01	——	.0009

Classes of organic foods. Our organic foods may be divided into three main classes: carbohydrates, fats, and protein. We get the carbohydrates in the form of starch and sugar from potatoes, corn, beans, bread, and many other sources. In the body, the carbohydrates are burned to furnish heat and power. If an excess is used, the body has the power to turn it into fat, and this is stored in the body.

We obtain fats in the form of butter, oils, fat of meat, and nuts. Most vegetables contain little or no fat. In the body, fats, like carbohydrates, are burned to furnish heat and energy. They are a more concentrated food, for one pound of fat is roughly equivalent to two and a quarter pounds of carbohydrates in available energy. The normal human body contains about 15% of fat. If more fat is eaten than can be oxidized, the excess may be stored in the body.

[1] From a table compiled by Henry C. Sherman.

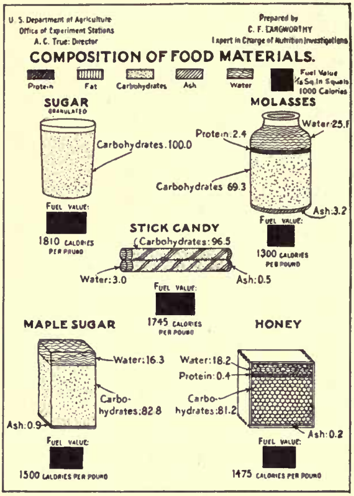


FIG. 90. — Composition of sugar products.

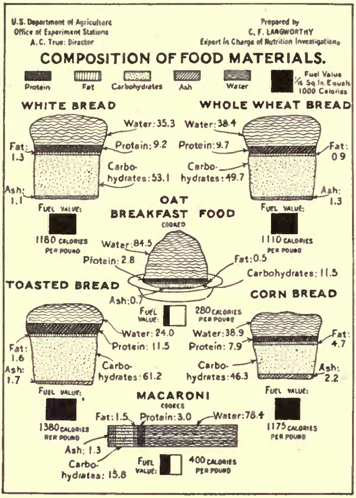


FIG. 91. — Composition of bread.

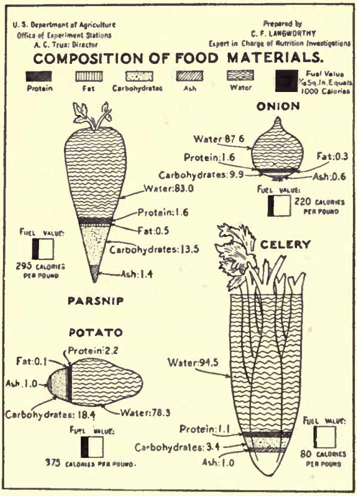


FIG. 92. — Composition of vegetables.

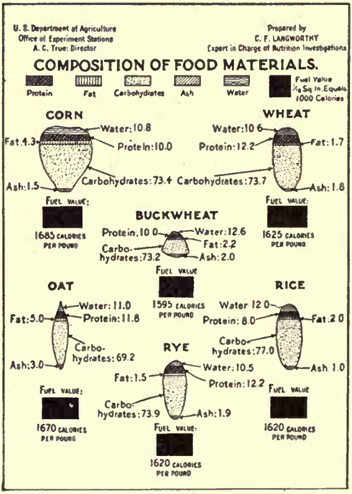


FIG. 93. — Composition of cereals.

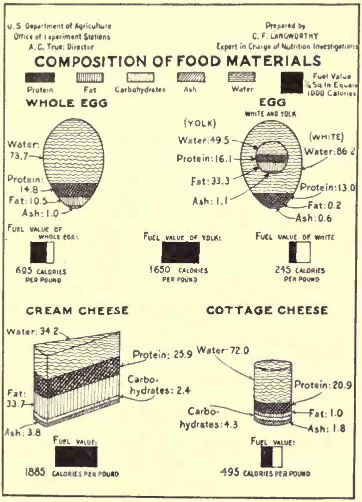


FIG. 94. — Composition of eggs and cheese.

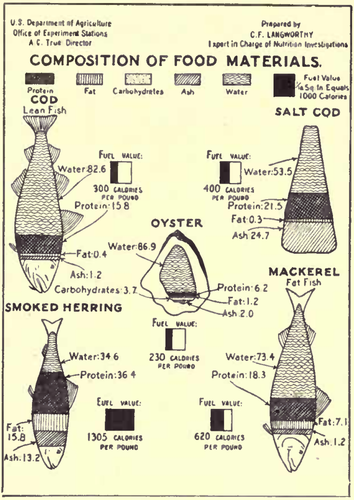


FIG. 95. — Composition of fish.

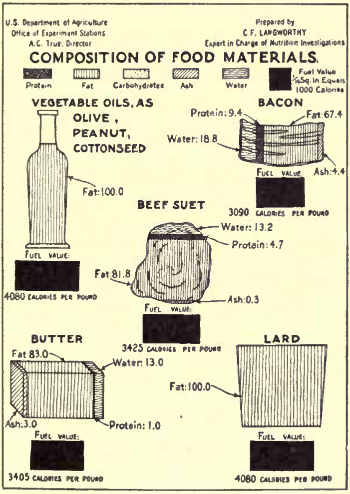


FIG. 96. — Composition of fats.

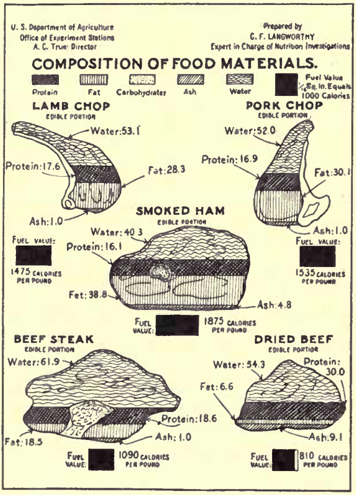


FIG. 97. — Composition of meats.

We get protein from the white of egg, from lean meat, and such vegetable foods as peas, beans, and the gluten of flour. The function of protein in the body is to build and repair the tissues, and it is the only food that can do this. It can also be burned in the body to furnish heat. This is not economical, as both fats and carbohydrates are cheaper sources of heat.

Analogy of the motor. In some ways our bodies resemble the gasoline motor that drives a car. If we wish our motor car to run properly, we must see that it has the proper fuel and use this in the right amount. If we use too much fuel, our car does not run well, and much of the fuel escapes unburned. The same thing is true of our bodies. If we overload our stomachs, we are dull and logy, and cannot do our work properly.

If we wish our car to run properly, we must carefully adjust the proportions of air and gasoline, and see that all the bearings are properly oiled. So with our bodies, if we wish them to run well, we must use all of the different foods in the proper proportions, and not limit our diet to any one class.

If we are using our car to carry heavy loads up steep hills, we must feed more gasoline to the motor. So too, if we are doing heavy manual work, our bodies require more fuel-food and also a slight increase in building-food, than if we are doing office work all day long.

One great difference between a car and our bodies is found in the result which follows neglect. If we abuse a car by neglecting to oil it properly, the car stops, and we have a heavy repair bill to pay. We may abuse our bodies, and for a while the body will struggle along, but in the end the result is the same. We have a heavy doctor's bill to pay, but unfortunately the physician cannot put a new part in our bodies, as the mechanic can in a car. He can only patch them up, and consequently we may be inefficient as long as we live.

A tire that will give 5000 miles of service under ordinary conditions will give but 200 miles when used in a race. Our bodies, too, are meant to be used moderately. If we burn the candle at both ends, it can be expected to last only half as long.

Of course our body is much more than a machine, for we have the power to think, and intellectual and spiritual powers that no machine has. But, even in this field, if we wish to develop our powers to the utmost, we must remember that the first requirement is a sound body, and that a sound body is best obtained by the use of proper foods in correct amounts.

Efficiency and economy. Large manufacturers do not buy coal solely by the ton. They have the coal analyzed, and its power to give energy determined. They do not rely altogether on price. Some coal at $4.00 a ton is cheaper than other coal at $3.00 a ton, for it contains more heat units; that is, when burned it will give much more heat than the coal sold at the lower price per ton. Its appearance will not tell us this; it must be analyzed.

The same thing is true of foods. It does not follow that the appearance and taste of food is an indication of its nutritive value. To determine its value to us, we must analyze it and find the quantities and proportions of the different food constituents it contains, and the amount of available energy it will yield to the body. We must also know the requirements of the body, and we are then in a position to provide wholesome food in the proper amounts. By knowing the food value of different foods in the market, we shall be able to buy nourishment economically.

The two charts (Figs. 98A and 98B) are intended to give an idea of the economic values of food from two standpoints. No. 1 considers the question from the standpoint of *Building Material;* No. 2 takes into account the cost of the *Energy Material,* or stuffs found in the different foods. Since most foods provide for us both building materials and energy, these charts should be studied together. It would be incorrect to attempt to get results from either chart to the exclusion of the other.

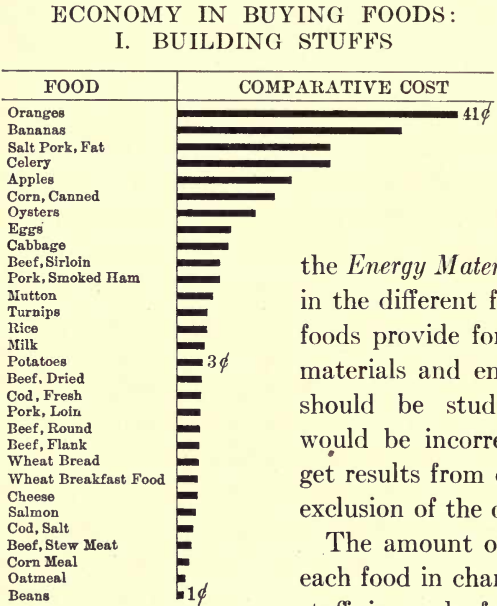

ECONOMY IN BUYING FOODS:
I. BUILDING STUFFS

Fig. 98A. — The cost of the same amount of building stuffs in different forms of foods.

The amount of building stuff in each food in chart No. 1, or energy stuff in each food in chart No. 2, is the same. The only variations shown in these charts are in the *prices* which we pay for these materials, as represented by the black bars of different lengths.

A little perusal will bring out the fact that the foods found nearest the top of each chart are the most expensive and the more economical ones are found toward the bottom of the list.

Although the number and range of foods in these small tables is necessarily limited, the lesson need be no less impressive. It must be remembered that only the financial side of the question is here considered and that digestion and personal idiosyncrasies are individual problems; it is believed, however, that a family that cares to live economically, can get enough in variety and substance in the lower two thirds of these charts, to nourish them palatably, economically, and well.

ECONOMY IN BUYING FOODS:
II. ENERGY STUFFS

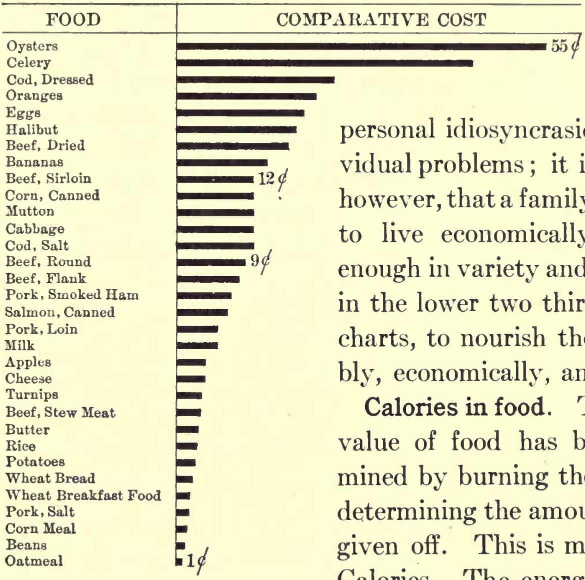


FIG. 98B. — The cost of the same amount of energy stuffs in different forms of foods.

Calories in food. The energy value of food has been determined by burning the food and determining the amount of heat given off. This is measured in Calories. The energy that the food will give when oxidized in the body is thus determined.

Respiration calorimeter. To study the requirements of the body under different conditions, the respiration calorimeter is used. One of the most famous of these is the one at Wesleyan University, built by Professors Atwater and Rosa. It is a copper-lined box, 7 feet long, 4 feet wide, and 6 feet 4 inches high. It is so arranged that a man can live in it for days, either resting, or doing work. All of the necessary food can be passed in to him, the heat given off by the body measured, the amount of work done measured, and the air going in and coming out analyzed. It is provided with a telephone, so that the subject can communicate with the outside without opening the box.

Men have lived in this box for from 3 to 12 days, and during all this time accurate analyses have been made of all food and air used, all work done has been measured, and all changes in the weight of the body determined. The effect of different diets, and the amount of food necessary to keep the body in condition, have been determined.

The problem has been attacked, in other ways, in many countries, until we can now estimate closely the amount of food compounds required to maintain the body under different conditions, and how these compounds may be obtained from different combinations of foods.

Balanced ration. A diet containing the proper proportions of fats, carbohydrates, and protein, to maintain the body in health, is called a balanced ration. In the army and navy, and in all large institutions, an effort is made to provide such a diet.

You will find in the tables beginning on page 261 the analyses of the different foods which we eat. From these you can calculate whether your diet is properly balanced.

According to recent experiments by Professor Chittenden, of Yale University, the ordinary man weighing 160 pounds, and doing a moderate amount of work, requires 2 oz. of available protein per day and enough of the fuel foods (fats and carbohydrates) to make the total fuel value of the food consumed 2500–3000 Calories.

Another way of stating this fact is: $\frac{1}{80}$ of an ounce of protein is required for each pound a person weighs, and enough of the fuel foods in addition to make up a total of 2500–3000 Calories per day.

The food requirements of the body vary with the amount of work done. The energy requirement which should be met with fats and carbohydrates varies almost proportionally with the amount of work done, while the protein requirement remains nearly constant. These statements apply to adults.

Disadvantage of excessive protein. The great fault with our American dietary is that it contains too much protein. Protein eaten, but not required by the body, is mainly broken down and excreted by the kidneys as uric acid and other poisonous substances. This is not only a waste, but throws too much of a strain on the kidneys, and leads to various diseases.

Ration for growing boys and girls. A growing child requires more food in proportion to his weight than an adult, as is explained by Frank A. Rexford in the text of his " A One-Portion Food Table."

" Boys and girls of school age present many difficulties in the way of rational feeding. Their needs are more in proportion to their size and weight than those of adults. They must have building stuff (protein) enough to keep them in repair, and in addition to this they must have a goodly amount of this material to grow on, for, if they are normal and healthy, they are continuously adding new tissues to their bodies. They must also have a generous supply of energy stuffs because of their many activities. We allow the average growing child 2 to 2½ ounces of protein per day and enough of the fuel stuffs (fat and carbohydrates) to bring the fuel value to 2000–3000 Calories according to his activity. A boy on the football team requires more energy stuff than the girl whose chief exercise consists in playing the piano. The tendency of most children is to eat nearly double the amount of protein they need and rather less of the fat and carbohydrates."

Condiments. There are certain adjuncts to a meal that are widely used, but contain little nutriment. These are condiments and spices, as mustard, pepper, catsup, and sauces. They may be useful, however, because they stimulate a jaded appetite, and increase the secretion of the digestive juices. The beverages, as tea and coffee, act as stimulants, but have no nutritive value. The use of all these is best kept at a minimum.

Use of the one-portion food table. Knowing the requirements of your body, you can use the food tables[1] on the following pages to plan a balanced dietary that will give you the proper amount of protein, and Calories enough to maintain your body at its highest efficiency. If you will note down for a few days the amount of food that you eat, you can readily, by using this table, determine whether your diet is a suitable one. If it is not, you will do well to correct it.

On page 271 is given a child's dietary taken at random from several thousands. The figures giving the values of the different " food stuffs " and the fuel value were copied from the table and totaled. The results show that the child was getting 3.77 ounces of protein per day and enough fat and carbohydrates to make the fuel reach 2890 Calories. In this particular case the child was instructed to lower

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[1] These food tables were computed and arranged in a convenient form by Mr. Rexford. He has worked for several years with the students of the Erasmus Hall High School and determined the weight of the ordinary helping of the common foods.

the protein total to 2.5 ounces and the fuel total to something between 2000 and 2500 Calories without diminishing the number of dishes. The modified dietary is shown on page 271.

FOODS PRIMARILY OF PLANT ORIGIN

Food as we eat It	Weight of Ordinary Helping	Of this the Body can use			This Portion can yield to the Body in Energy and Heat Units
		Muscle Builder	For Heat and Energy		
		Protein	Fat	Carbohydrates (Starch and Sugar)	
Beverages	Ounces	Ounces	Ounces	Ounces	Calories
Cocoa	5.	.11	.33	.19	123.
Coffee (cream and sugar only)	.75	.01	.17	.27	53.
Lemonade	5.5	—	—	.66	78.1
Orange juice	5.	—	—	.65	75.5
Bread					
Biscuit, cream	2.33	.2	.2	1.	203.9
home made	2.	.17	.05	1.1	162.5
soda	2.	.19	.27	1.05	216.3
Bread, corn	2.	.16	.09	.93	150.6
gluten	2.	.18	.03	.99	145.
graham	2.	.18	.04	1.04	151.3
home made	2.	.18	.03	1.07	153.1
plain rolls	2.	.19	.08	1.2	182.08
rye	2.	.23	.01	.71	148.4
whole wheat	2.	.19	.02	.99	142.5
zwieback	1.	.1	.1	.74	123.2
and butter	2.5	.22	.48	1.18	275.
Buns, hot cross	1.25	.1	.06	.6	99.6
Crackers,					
graham	1.	.1	.09	.74	122.2
oatmeal	1.	.12	.11	.69	123.1
pretzels	1.	.1	.04	.73	106.3
saltines	1.	.11	.13	.69	125.3
soda	1.	.1	.09	.73	120.3
Toast, cream	5.	.2	.56	.6	238.5
Toast, dry	5.	.06	.008	.3	44.4

FOODS PRIMARILY OF PLANT ORIGIN — *Continued*

Food as we eat It	Weight of Ordinary Helping	Of this the Body can use			This Portion can yield to the Body in Energy and Heat Units
		Muscle Builder	For Heat and Energy		
		Protein	Fat	Carbohydrates (Starch and Sugar)	
Cake	Ounces	Ounces	Ounces	Ounces	Calories
Chocolate, layer . . .	2.5	.14	.2	1.6	256.8
Charlotte russe	4.25	.26	.56	2.39	395.2
Coffee . . .	2.	.14	.15	1.26	203.1
Cookies,					
molasses .	1.75	.13	.16	1.32	209.
sugar . . .	1.5	.11	.15	1.1	180.
Doughnuts .	1.75	.12	.37	.93	218.8
Frosted . . .	2.	.12	.18	1.3	211.9
Fruit . . .	2.	.12	.22	1.28	220.
Gingerbread .	2.	.12	.18	1.37	208.8
Jelly roll . .	3.	.15	.12	2.19	301.2
Lady fingers .	.5	.04	.03	.35	52.7
Macaroons .	1.	.07	.15	.65	123.4
Nut . . .	2.5	.2	.54	1.36	324.4
Sponge . . .	1.5	.09	.16	.9	168.3
Cereals					
Corn flakes .	.75	.07	.003	.59	77.6
Farina . . .	4.	.44	.06	2.98	421.2
Hominy . .	4.	.08	.01	.72	95.
Oatmeal . .	4.25	.13	.02	.49	76.5
Puffed rice .	.5	.04	—	.4	50.9
Rice	4.	.11	.001	.96	124.72
Shredded wheat (2) .	2.	.21	.03	1.56	212.5
Wheat flakes .	.75	.1	.01	.56	79.2
Fruit					
Apple, baked .	3.25	.02	.02	.78	98.5
Apple, fresh .	5.5	.02	.02	.78	99.6
Apple, sauce .	3.5	.01	.03	1.3	76.1
Bananas . .	3.5	.05	.02	.77	100.8

FOODS PRIMARILY OF PLANT ORIGIN — *Continued*

Food as we eat It	Weight of Ordinary Helping	Of this the Body can use			This Portion can yield to the Body in Energy and Heat Units
		Muscle Builder	For Heat and Energy		
		Protein	Fat	Carbohydrates (Starch and Sugar)	
Fruit (*cont'd*)	Ounces	Ounces	Ounces	Ounces	Calories
Cherries . .	2.	.03	.01	.44	45.6
Currants . .	3.	.04	—	.37	50.1
Cranberries .	3.	.01	.02	.3	40.3
Dates . . .	1.75	.04	.05	.59	177.6
Figs	2.	.09	.01	1.5	184.4
Grapefruit .	3.75	.03	.01	.37	49.5
Grapes . . .	5.	.05	.06	.71	104.5
Huckleberries .	3.	.02	.02	.5	64.6
Lemons . .	1.	.01	—	.08	13.2
Olives, green .	1.33	.01	.37	.15	116.3
Olives, ripe .	1.33	.02	.33	.06	100.1
Oranges . .	5.	.04	.01	.58	75.
Peaches, cooked	3.5	.04	.04	.54	73.2
Pineapple,					
canned . .	3.25	.013	.227	1.18	145.21
fresh . . .	4.	.016	.012	.38	52.48
Prunes, cooked	3.75	.03	.004	.85	103.1
Raspberries,					
black . . .	4.	.07	.04	.5	77.2
red . . .	3.5	.04	—	.44	55.7
Rhubarb . .	2.5	.01	.01	.57	67.
Strawberries,					
fresh . . .	4.25	.04	.03	.31	48.5
Jelly					
Cherry . . .	1.	.01	—	.21	90.9
Cranberry . .	2.	.01	.01	.85	102.2
Currant . .	1.	.01	—	.77	91.3
Orange . . .	2.75	—	—	.85	100.1
Peach . . .	3.5	.02	.05	.74	98.4
Miscellaneous					
Brown gravy .	2.25	.03	.26	.07	81.2
Hash, beef . .	2.2	.26	.27	.32	114.3

FOODS PRIMARILY OF PLANT ORIGIN — *Continued*

FOOD AS WE EAT IT	WEIGHT OF ORDINARY HELPING	OF THIS THE BODY CAN USE			THIS PORTION CAN YIELD TO THE BODY IN ENERGY AND HEAT UNITS
		Muscle Builder	For Heat and Energy		
		Protein	Fat	Carbohydrates (Starch and Sugar)	
Miscellaneous (*cont'd*)	Ounces	Ounces	Ounces	Ounces	Calories
Macaroni . .	2.75	.36	.02	2.00	286.2
Macaroni with cheese . .	2.75	.26	.16	.42	122.4
Mayonnaise dressing (cooked) .	1.25	.07	.32	.03	96.
Olive oil (tablespoon)	.33	—	.33	—	88.
Salad dressing (French) .	.5	—	.74	.02	100.4
Nuts					
Almonds . .	.25	.05	.14	.04	47.8
Beech . . .	.5	.11	.29	.07	97.
Brazil . . .	.5	.09	.33	.04	103.
Butter . . .	.5	.14	.31	.02	99.9
English walnuts . .	.5	.08	.32	.08	103.4
Filberts . .	.5	.08	.33	.07	103.7
Hickory . .	.5	.08	.34	.06	105.5
Peanuts . .	.5	.13	.19	.12	80.1
Pecan . . .	.5	.06	.36	.7	108.9
Pickles					
Cucumber .	1.25	.006	.004	.03	5.4
Mixed . . .	1.	.01	.004	.04	6.9
Spiced . . .	1.	.004	.001	.21	24.7
Pie					
Apple . . .	4.5	.29	.31	1.44	282.8
Blueberry . .	3.87	.15	.19	1.5	237.
Cream . . .	4.	.18	.46	2.05	380.
Custard . .	4.	.17	.25	1.	207.2

FOODS PRIMARILY OF PLANT ORIGIN—*Continued*

Food as we eat It	Weight of Ordinary Helping	Of this the Body can use			This Portion can yield to the Body in Energy and Heat Units
		Muscle Builder	For Heat and Energy		
		Protein	Fat	Carbohydrates (Starch and Sugar)	
Pie (*cont'd*)	Ounces	Ounces	Ounces	Ounces	Calories
Coconut cream . .	3.87	.23	.46	1.04	235.2
Lemon . . .	4.	.14	.4	1.4	297.2
Mince . . .	5.	.65	.42	1.51	362.
Pumpkin . .	5.	.15	.15	1.	177.
Raisin . . .	5.	.15	.56	2.36	439.5
Squash . . .	5.	.22	.42	1.08	265.5
Pudding					
Blanc mange (chocolate)	3.5	.1	.3	.49	148.8
Bread . . .	3.5	.19	.42	.57	131.6
Custard . .	3.25	.16	.16	.35	102.4
Date . . .	2.5	.15	.23	1.4	243.
Fig	2.75	.11	.17	.82	150.4
Floating island	3.	.15	.05	.55	118.8
Indian meal .	3.25	.18	.16	.89	165.5
Rice . . .	3.25	.12	.28	.55	149.5
Snow . . .	2.5	.1	.07	.35	75.9
Tapioca . .	3.25	.11	.1	.92	146.3
and apple .	3.25	.01	.003	.95	112.4
Salad					
Date and apple	2.25	.05	.05	.87	121.7
Date and walnut . .	1.25	.63	.16	.62	124.1
Egg mayonnaise . . .	2.25	.26	.25	.02	100.1
Fruit . . .	2.25	.04	.02	.52	70.4
Potato . . .	2.25	.09	.22	.29	102.1
String bean .	1.75	.01	.33	.04	95.7
Tomato (with mayonnaise)	4.	.06	.16	.15	67.6

FOODS PRIMARILY OF PLANT ORIGIN — *Continued*

Food as we eat It	Weight of Ordinary Helping	Of this the Body can use			This Portion can yield to the Body in Energy and Heat Units
		Muscle Builder	For Heat and Energy		
		Protein	Fat	Carbohydrates (Starch and Sugar	
Soup	Ounces	Ounces	Ounces	Ounces	Calories
Bean . . .	4.75	.38	.07	1.	182.8
Cream of					
celery . .	4.75	.11	.34	.17	124.8
corn . . .	4.75	.14	.33	.4	152.
Chicken . .	4.75	.19	.01	.1	31.3
Consommé .	4.75	.1	—	.02	16.
Clam chowder	4.75	.09	.04	.33	60.
Lentil . . .	4.75	.23	.25	.48	161.5
Oxtail . . .	4.75	.2	.06	.2	65.
Potato . . .	4.75	.11	.03	.37	146.8
Tomato . .	4.75	.13	.12	.33	91.2
Vegetable (canned) . .	4.75	.13	—	.02	192.8
Sugars					
Candy,					
caramel . .	1.	.05	—	.81	100.4
chocolate .	1.	.01	.01	.73	90.
Chocolate . . almonds .	1.5	.06	.15	.95	160.
Honey . . .	1.63	.1	—	1.32	155.2
Maple sirup .	1.25	—	—	.89	103.9
Maple sugar .	1.	—	—	.83	96.6
Sugar (granulated or loaf)	.25	—	—	.25	27.
Vegetables					
Asparagus (on toast) . .	4.	.18	.4	.64	202.8
Beans, baked .	3.25	.31	.18	1.08	182.
kidney . .	3.25	.22	.65	.6	97.5
string . .	4.	.69	.61	.21	48.72
Beets . . .	2.25	.05	.002	.15	26.1

FOODS PRIMARILY OF PLANT ORIGIN — *Continued*

FOOD AS WE EAT IT	WEIGHT OF ORDINARY HELPING	OF THIS THE BODY CAN USE			THIS PORTION CAN YIELD TO THE BODY IN ENERGY AND HEAT UNITS
		Muscle Builder	For Heat and Energy		
		Protein	Fat	Carbohydrates (Starch and Sugar)	
Vegetables (*cont'd*)	Ounces	Ounces	Ounces	Ounces	Calories
Cabbage, boiled	4.	.03	.09	.16	35.2
Carrots . .	3.75	.04	.02	.35	49.2
Cauliflower .	4.	.07	.02	.2	35.
Celery . . .	1.	.01	.003	.04	5.5
Corn, canned	2.75	.08	.03	.52	74.25
Cucumbers .	2.	.03	.004	.06	10.
Egg plant . .	1.5	.09	.15	.48	106.5
Lettuce . .	1.	.01	—	.03	7.
Mushrooms .	1.	.04	.01	.07	13.1
Onions, boiled	2.5	.03	.11	.13	26.3
creamed .	3.	.04	.15	.15	65.7
scalloped .	3.	.08	.27	.09	73.9
Parsnips,					
creamed .	3.	.03	.07	.44	79.
browned .	3.	.05	.13	.29	75.6
Peas, canned .	3.	.09	.09	.54	66.6
green . .	3.	.13	.1	.2	103.2
Potatoes, sweet	3.	.09	.06	1.26	173.4
Potatoes,					
baked . .	3.	.1	.01	.68	98.1
boiled . .	3.	.08	.01	.73	82.8
browned .	3.25	.11	.06	.82	123.5
mashed . .	3.25	.09	.26	.68	100.4
Radishes . .	1.	.01	—	.05	8.5
Spinach . .	3.	.06	.12	.09	16.3
Squash, winter					
(baked) . .	3.75	.05	.01	.33	50.4
Succotash . .	3.	.11	.03	.56	78.
Tomatoes,					
sliced . .	4.	.04	.04	.18	26.8
stewed . .	2.5	.03	—	.08	16.4
Turnips,					
mashed . .	4.	.02	.11	.11	24.4

FOODS PRIMARILY OF ANIMAL ORIGIN

Food as we eat It	Weight of Ordinary Helping	Of this the Body can use			This Portion can yield to the Body in Energy and Heat Units
		Muscle Builder	For Heat and Energy		
		Protein	Fat	Carbohydrates (Starch and Sugar)	
Beef	Ounces	Ounces	Ounces	Ounces	Calories
Chuck . . .	3.	.57	.04	—	172.5
Corned . . .	2.	.21	.52	—	174.2
Dried . . .	1.	.26	.07	—	49.4
Flank . . .	2.25	.44	.47	—	176.4
Heart . . .	1.	.16	.2	—	72.5
Liver . . .	2.	.41	.09	.03	75.6
Round . . .	2.25	.43	.29	—	125.2
Sirloin . . .	2.25	.37	.36	—	137.1
Sweetbreads .	2.	.33	.24	—	103.1
Tongue, pickled . .	2.	.21	.41	—	138.
Tripe . . .	3.	.38	.04	.01	60.4
Dairy Products					
Butter . . .	.5	.05	.43	—	112.5
Buttermilk .	6.	.18	.03	.29	61.9
Cheese, cottage . .	2.	.31	.09	.09	74.6
full cream .	1.	.26	.34	.02	122.4
Neuchatel .	2.	.37	.55	.03	191.3
pineapple .	2.	.6	.78	.05	280.6
Swiss . . .	1.	.22	.35	.01	125.6
Condensed milk (sweetened)	.25	.02	.02	.14	23.8
(unsweetened) . .	.25	.02	.02	.03	10.6
Cream, (tablespoon)	.5	.01	.17	.02	26.
whipped .	.5	.13	.09	.05	31.1
Ice cream . .	2.	.05	.1	.91	134.7
Milk, skimmed	6.5	.22	.02	.33	71.3
whole . .	6.	.19	.24	.3	123.6
Oleomargarine	.5	—	.4	—	110.2

FOODS PRIMARILY OF ANIMAL ORIGIN — *Continued*

Food as we eat It	Weight of Ordinary Helping	Of this the Body can use			This Portion can yield to the Body in Energy and Heat Units
		Muscle Builder	For Heat and Energy		
		Protein	Fat	Carbohydrates (Starch and Sugar	
Eggs	Ounces	Ounces	Ounces	Ounces	Calories
Boiled (2)	3.75	.49	.45	—	179.1
Omelet	4.	.48	.88	.03	296.
Poached	1.25	.18	.15	—	60.4
on toast	2.5	.3	.17	.29	144.4
Scrambled	2.	.24	.17	.03	78.5
Uncooked (2)	3.75	.49	.45	—	179.1
Fish					
Blue	5.	1.3	.23	—	209.4
Cod	5.	.32	.02	—	101.6
Halibut, steak	3.	.56	.16	—	105.9
Herring	.5	.56	.16	—	105.9
Salmon, canned	2.	.44	.24	—	114.1
Sardines	1.	.23	.19	—	78.1
Shad	2.25	.42	.22	—	104.9
Trout (brook)	1.75	.33	.36	—	135.9
Fowl					
Chicken					
(broilers)	3.5	.75	.09	—	110.5
fricasseed	3.5	.62	.4	.08	187.
Goose	2.75	.43	.98	—	312.1
Turkey	1.25	.26	.26	—	104.
Lamb					
Chops	2.	.43	.59	—	210.
Kidney stew	4.	.72	.2	.08	150.
Leg	3.5	.67	.44	—	194.3
Mutton					
Leg	2.5	.62	.51	—	108.

FOODS PRIMARILY OF ANIMAL ORIGIN—*Continued*

Food as we eat It	Weight of Ordinary Helping	Of this the Body can use			This Portion can yield to the Body in Energy and Heat Units
		Muscle Builder	For Heat and Energy		
		Protein	Fat	Carbohydrates (Starch and Sugar)	
Pork	Ounces	Ounces	Ounces	Ounces	Calories
Bacon . . .	1.	.1	.66	—	188.6
Chops . . .	3.	.47	.95	—	309.
Ham croquettes	2.	.3	.24	.11	111.2
Ham, lean . .	2.25	.49	.55	—	203.2
Sandwiches					
Cheese . . .	3.25	.41	.49	1.2	314.2
Egg	4.	.4	.37	1.19	279.7
Ham . . .	3.5	.33	.48	1.19	302.7
Jelly . . .	2.	.01	.01	.85	102.2
Lamb . . .	4.	.53	.48	1.1	316.
Lettuce and mayonnaise	3.5	.22	.24	1.21	214.2
Roast beef .	3.5	.39	.39	1.19	275.8
Sardine . .	3.5	.44	.35	1.19	279.6
Sausages					
Bologna . .	2.	.37	.35	.01	136.8
City . . .	2.	.35	.48	—	172.5
Country . .	2.	.56	.8	—	278.1
Frankfurters .	2.	.39	.37	.02	141.6
Shellfish					
Clams	3.75	.24	.02	—	32.2
Lobster . .	2.	.32	.04	—	47.6
Oysters . .	3.5	.21	.04	—	36.4
Scallops . .	2.	.29	—	.07	43.1
Veal					
Breast (lean)	2.5	.38	.25	—	104.8
Cutlets . . .	3.5	.7	.26	—	152.
Leg	2.5	.65	.1	—	104.
Liver . . .	3.	.56	.17	—	107.1

ORDINARY DAILY DIETARY

Food for One Day	Of this the Body can use			This Portion can yield to the Body in Energy and Heat Units
	Protein	Fat	Carbohydrates (Starch and Sugar)	
Breakfast				
Oranges	.04	.01	.58	75.
Coffee	.01	.17	.27	53.
Bread and butter	.22	.48	1.18	275.
Omelet	.48	.88	.03	296.
Lunch				
3 Beef sandwiches	1.17	1.17	3.57	827.4
Orange	.04	.01	.58	75.
Chocolate candy	.01	.01	.73	90.
Dinner				
Lamb chops	.43	.59	—	210.
Potatoes, mashed	.09	.26	.68	103.4
Peas	.09	.09	.54	66.6
Lettuce	.01	—	.03	7.
Pineapple cheese	.60	.78	.09	280.6
Crackers, soda	.10	.09	.73	120.3
Apple pie	.29	.31	1.44	287.8
Milk	.19	.24	.30	123.6
Total	3.77	5.09	10.75	2890.7

MODIFIED DIETARY

Breakfast				
Oranges	.04	.01	.58	75.
Rice	.11	.001	.96	124.72
Cocoa	.11	.33	.19	123.
Dry toast	.06	.008	.3	44.4
Butter	.05	.43	—	112.5
Lunch				
Bean soup	.38	.07	1.	182.8
Lettuce and mayonnaise sandwiches	.44	.48	2.42	428.4
Dates	.04	.05	.59	177.6

MODIFIED DIETARY — *Continued*

Food for One Day	Of this the Body can use			This Portion can yield to the Body in Energy and Heat Units
	Protein	Fat	Carbohydrates (Starch and Sugar)	
Dinner				
Cream of celery soup	.11	.34	.17	124.8
Breast of veal . . .	.38	.25	—	104.8
Baked potatoes . .	.1	.01	.26	98.1
Green peas	.13	.1	.2	103.2
Salad, string bean .	.01	.33	.04	95.7
Bread and butter . .	.22	.48	1.18	275.
Pudding, floating island	.15	.05	.55	118.8
Whole milk	.19	.24	.3	123.6
3 O'Clock Lunch				
Lemonade	—	—	.66	78.1
Total	2.52	3.179	9.40	2390.52

Family food table.[1] Another food table by Mr. Rexford refers to the analysis of food as bought in the market. By keeping the family market lists for a week, and computing the nourishment in the foods bought by this table, you can compute the actual nourishment obtained for the money expended. The use of this table will lead to economy in buying foods.

In the study of this table a knowledge of the use and value of each of the different foodstuffs, as given in "A One-Portion Food Table," is presupposed. Up to this time our study of nutrition has been confined to the consideration of the needs of the individual and of *food as eaten*. When we come

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[1] Separate copies of Rexford's food tables, classroom and home record blanks, can be obtained from Educational Equipment Co., 70 Fifth Ave., New York, N. Y.

to *buying food* the proposition is a little different. We do not buy food by the portion; we buy it as it is sold in the market. The analyses here given are therefore in most cases computations from government bulletins and are for quantities of different foods in just the condition we buy them. That is, meat with the bone, lobster with the shell, peas with the pods, nuts with the shell, etc.

The following family market list shows the food bought for seven people for one week. Work it out, and answer the following questions:

1. Is this family of seven people getting the proper amount of protein?
2. Is its fuel value well regulated?
3. Is this an economical or an extravagant weekly list?

FAMILY MARKET LIST

3½ lb.	pork, loin	@	24c	1 head lettuce	@	10c
3 lb	lamb, leg	"	22c	3 baskets potatoes	"	20c
½ lb.	dried beef	"	25c	2 qt. onions	"	15c
1 lb.	codfish, salt	"	18c	3 bunches celery	"	12c
2½ lb.	beef, sirloin	"	28c	6 qt. spinach	"	15c
1 lb.	bacon	"	26c	1 hd. cabbage	"	6c
1 lb.	sausage	"	20c	3 cans peaches	"	25c
10 lb.	beef, roast	"	25c	2 doz. bananas	"	15c
10 loaves	bread, white	"	10c	4 qt. apples	"	16c
5 loaves	bread, graham	"	10c	1 doz. oranges	"	40c
4 doz.	rolls	"	10c	1 lb. coffee	"	30c
1 pkg.	oatmeal	"	10c	¼ lb. tea	"	50c
2 lb.	crackers, soda	"	15c	1 can cocoa	"	40c
1 pkg.	tapioca	"	10c	1 lb. walnuts	"	23c
7 lb.	sugar, gran.	"	36c	4 lb. butter	"	37c
21 qt.	milk	"	9c	4 doz. eggs	"	40c

NOTE. — It is not necessary to total the fat and carbohydrate columns, as their value, in the main, is given in the Calorie column. It is well to record them, however, in order to keep in mind the proportions of fuel ingredients found in the different foods, so that the fuel value may be controlled by increasing or curtailing them, as the case demands.

This work has value only when it is applied. It is desirable that each family in the community work out a weekly market list and compare notes. This can be done by keeping account in each case of what is bought and the price paid. A tabulation of the results will show: (*a*) comparative cost of the same foods in different parts of the town; (*b*) whether or not the family is being approximately well nourished; (*c*) whether or not the family is getting good value in *nourishment* for the money expended each week.

FAMILY FOOD TABLE

Foods Primarily of Plant Origin

Food as we buy It	Nourishment in This Food			This Food can yield to the Body in Energy and Heat Units
	Muscle Builder	For Heat and Energy		
	Protein	Fat	Carbohydrates (Starch and Sugar)	
Beverages	Pounds	Pounds	Pounds	Calories
Cocoa, 1 lb. can . . .	.22	.28	.38	2320
Coffee, 1 lb.	.00	.00	.00	0000
Tea, 1 lb.	.00	.00	.00	0000
Bread				
Buns, 1 doz., 1 lb. . .	.07	.08	.58	1515
Graham, 5c. loaf, 1 lb. .	.06	.01	.39	908
Rolls, 1 doz., 1 lb. . .	.10	.04	.60	1470
Rye, 5c. loaf, 12 oz. . .	.07	.01	.40	912
White, 5c. loaf, 12 oz. .	.07	.01	.40	912
Whole Wheat, 12 oz. loaf	.08	.01	.37	855
Crackers				
Graham, 1 lb.	.10	.09	.73	1955
Oatmeal, 1 lb.	.12	.11	.69	1970
Pretzels, 1 lb.	.10	.04	.73	1700
Saltines, 1 lb.	.11	.13	.69	2005
Soda, 1 lb.	.10	.09	.73	1925

FOODS PRIMARILY OF PLANT ORIGIN — *Continued*

FOOD AS WE BUY IT	NOURISHMENT IN THIS FOOD			THIS FOOD CAN YIELD TO THE BODY IN ENERGY AND HEAT UNITS
	Muscle Builder	For Heat and Energy		
	Protein	Fat	Carbohydrates (Starch and Sugar)	
Cake	Pounds	Pounds	Pounds	Calories
Charlotte russe, 1, 4.25 oz.	.02	.04	.17	395
Chocolate, layer, $3\frac{1}{3}$ lb.	.20	.27	2.10	5500
Cookies, molasses, 1 doz., 12 oz.	.07	.15	.65	1975
Cookies, sugar, 1 doz., 12 oz.	.05	.07	.55	1440
Doughnuts, 1 doz., 12 oz.	.05	.16	.40	1500
Fruit, 1 lb.	.06	.11	.64	1760
Gingerbread, 1 lb.	.06	.09	.63	1670
Lady fingers, 1 doz., 4 oz.	.02	.02	.18	422
Macaroons, 1 lb.	.07	.15	.65	1975
Sponge, 1 lb.	.06	.11	.66	1795
Cereals				
Farina, 1 box, 28 oz.	.19	.02	1.34	2949
Hominy, 1 box, 28 oz.	.15	.01	1.38	2888
Oatmeal, 1 box, 28 oz.	.22	.10	.93	2558
Rice, 1 lb.	.08	.00	.79	1630
Shredded wheat, 1 box, 12 oz.	.09	.01	.56	1275
Wheat breakfast foods, 1 lb.	.12	.02	.75	1680
Flours, etc.				
Buckwheat, 1 bag, $3\frac{1}{2}$ lb.	.22	.04	.27	5607
Corn meal, 1 bag, $3\frac{1}{2}$ lb.	.32	.07	2.54	5722
Graham, 1 bag, $3\frac{1}{2}$ lb.	.45	.03	2.50	5775
Macaroni, 1 box, 12 oz.	.09	.01	.56	1233
Rye, 1 bag, $3\frac{1}{2}$ lb.	.24	.03	2.75	5705
Spaghetti, 12 oz. box	.09	.00	.57	1245
Wheat, 1 bag, $3\frac{1}{2}$ lb	.38	.04	2.62	5740
Wheat, 1 bag, 7 lb.	.76	.08	5.24	11480
Wheat, 1 bag, $24\frac{1}{2}$ lb	2.65	.27	18.32	40180
Whole wheat, 1 bag, 5 lb.	.66	.09	3.24	7956

Foods Primarily of Plant Origin — *Continued*

Food as we buy It	Nourishment in This Food			This Food can yield to the Body in Energy and Heat Units
	Muscle Builder	For Heat and Energy		
	Protein	Fat	Carbohydrates (Starch and Sugar)	
Fruit	Pounds	Pounds	Pounds	Calories
Apples, 1 qt., 28 oz.	.01	.01	.19	385
Bananas, 1 doz., 4½ lb.	.04	.02	.64	1350
Cherries, 1 can, 30 oz.	.02	.00	.40	778
Cranberries, 1 lb.	.01	.01	.10	215
Dates, 1 box, 12 oz.	.01	.02	.53	1088
Figs, 1 lb.	.04	.00	.74	1475
Grapes, 1 lb.	.01	.01	.14	335
Lemons, 1 doz., 3 lb.	.03	.02	.26	615
Oranges, 1 doz., 5¼ lb.	.04	.01	.60	1200
Peaches, 1 can, 28 oz.	.01	.00	.19	385
Pears, 1 can, 28 oz.	.01	.01	.31	621
Pineapple, 1 can, 28 oz.	.01	.01	.64	1251
Prunes, 1 lb.	.02	.00	.73	1400
Jellies (average analysis)				
1 jar, 10 oz.	.01	.00	.50	950
Miscellaneous				
Olive oil, 1 pt. bottle, 1 lb.	.00	1.00	.00	4226
Nuts				
Almonds, 1 lb.	.12	.30	.10	1660
Brazil nuts, 1 lb.	.09	.34	.04	1655
Filberts, 1 lb.	.08	.31	.06	1575
Hickory nuts, 1 lb.	.06	.26	.04	1265
Peanuts, 1 lb.	.20	.29	.19	1935
Peanut butter, 1 jar, ¼ lb.	.07	.12	.04	706
Pecans, 1 lb.	.05	.33	.06	1620
Walnuts, 1 lb.	.07	.27	.07	1375
Pickles				
Cucumber, 14½ oz. bottle.	.00	.00	.02	61
Mixed, 1 bottle, 14 oz.	.01	.00	.04	96
Spiced, 1 bottle, 14 oz.	.00	.00	.19	370

FOODS PRIMARILY OF PLANT ORIGIN — *Continued*

FOOD AS WE BUY IT	NOURISHMENT IN THIS FOOD			THIS FOOD CAN YIELD TO THE BODY IN ENERGY AND HEAT UNITS
	Muscle Builder	For Heat and Energy		
	Protein	Fat	Carbohydrates (Starch and Sugar)	
Pies	Pounds	Pounds	Pounds	Calories
Apple, 1, 2¼ lb. . . .	.07	.22	.96	2857
Custard, 1, 2¼ lb. . . .	.09	.14	.58	1867
Mince, 1, 2¼ lb. . . .	.14	.28	.86	3003
Soup				
Asparagus, cream of, 1 can, 1 lb.	.03	.03	.05	285
Bouillon, 1 can, 1 lb. .	.02	.00	.00	50
Celery, cream of, 1 can, 1 lb.	.02	.03	.05	235
Chicken, 1 can, 1 lb. .	.04	.00	.02	100
Consommé, 1 can, 1 lb.	.03	.00	.00	55
Corn, cream of, 1 can, 1 lb.	.03	.02	.08	270
Mock turtle, 1 can, 1 lb.	.05	.01	.03	185
Mulligatawny, 1 can, 1 lb.	.04	.00	.06	180
Oxtail, 1 can, 1 lb. . .	.04	.01	.04	170
Pea, 1 can, 1 lb. . . .	.04	.01	.08	235
Pea, cream of, 1 can, 1 lb.	.03	.03	.06	270
Tomato,	.02	.01	.06	185
Turtle, 1 can, 1 lb. . .	.06	.02	.04	265
Vegetable, 1 can, 1 lb. .	.03	.00	.01	65
Sugars and Starches				
Candy, 1 lb.	.00	.00	.96	1680
Cornstarch, 1 package, 1 lb.	.00	.00	.90	1675
Sugar, brown, 1 lb. . .	.00	.00	.95	1765
Sugar, gran., 1 bag, 3½ lb.	.00	.00	3.50	6510
Sugar, maple, 1 lb. . .	.00	.00	.82	1540
Sugar, powdered, 1 lb. .	.00	.00	1.	1860
Molasses, 1 can, 29 oz. .	.04	.00	1.25	2338
Tapioca, 1 package, 14 oz.	.00	.00	.77	210

Foods Primarily of Plant Origin — *Continued*

Food as we buy It	Nourishment in This Food			This Food can yield to the Body in Energy and Heat Units
	Muscle Builder	For Heat and Energy		
	Protein	Fat	Carbohydrates (Starch and Sugar)	
Vegetables	Pounds	Pounds	Pounds	Calories
Asparagus, 1 can, 28 oz.	.02	.00	.05	149
Beans, string, 1 qt.	.04	.01	.12	315
Beets, 1 bunch, 2 lb.	.02	.00	.15	340
Cabbage, 1 head, 4 lb.	.01	.00	.17	500
Carrots, 1 bunch, 10 oz.	.01	.00	.05	100
Cauliflower, 1 head, 2 lb.	.04	.01	.09	280
Celery, 1 bunch, 10 oz.	.00	.00	.01	44
Corn, 1 can, 19 oz.	.03	.01	.23	540
Cucumber, 1 large, 12 oz.	.01	.00	.02	51
Eggplant, 1 lb.	.01	.00	.05	130
Lettuce, 1 head, 10 oz.	.01	.00	.02	51
Mushrooms, 1 lb.	.04	.00	.07	210
Onions, 1 qt., 28 oz.	.03	.01	.16	359
Parsnips, 1 bunch, 24 oz.	.02	.01	.16	360
Peas, canned, 1 can, 20 oz.	.04	.01	.17	419
Peas, green, 1 qt., 28 oz.	.06	.00	.17	446
Potatoes, Irish, 1 basket, 6 lb.	.10	.01	.88	1860
Potatoes, sweet, 1 qt., 28 oz.	.02	.01	.38	805
Radishes, 1 bunch, 5 oz.	.00	.00	.01	30
Spinach, 1 qt., 28 oz.	.04	.01	.06	193
Succotash, 1 can, 19 oz.	.04	.01	.22	540
Tomatoes, 1 lb.	.01	.00	.04	105
Tomatoes, canned, 1 can, 38 oz.	.03	.01	.10	249
Turnips, 1, 3 lb.	.03	.00	.17	375

FOODS PRIMARILY OF ANIMAL ORIGIN

FOOD AS WE BUY IT	NOURISHMENT IN THIS FOOD			THIS FOOD CAN YIELD TO THE BODY IN ENERGY AND HEAT UNITS
	Muscle Builder	For Heat and Energy		
	Protein	Fat	Carbohydrates (Starch and Sugar)	
Beef	Pounds	Pounds	Pounds	Calories
Chuck, roast or steak, 1 lb.	.15	.00	.00	920
Corned, 1 lb.	.14	.22	.00	1195
Dried, 1 lb.	.26	.07	.00	780
Flank, roast or steak, 1 lb.	.19	.20	.00	1185
Liver, 1 lb.	.20	.03	.03	555
Round, roast or steak, 1 lb.	.19	.09	.00	745
Sirloin, roast or steak, 1 lb.	.17	.17	.00	985
Sweetbreads, 1 lb.	.17	.12	.00	825
Tongue, 1 lb.	.14	.07	.00	545
Tripe, 1 lb.	.12	.01	.00	270
Dairy Products				
Butter, 1 lb.	.00	.85	.00	3605
Buttermilk, 1 qt., 2 lb.	.05	.06	.01	330
Cheese, cottage, 1 lb.	.21	.01	.04	510
Cheese, American, 1 lb.	.26	.34	.02	1950
Cream, 1 bottle, ½ pt., 8 oz.	.01	.08	.02	430
Ice cream, 1 qt., 2 lb.	.06	.18	.44	1808
Milk, 1 qt., 2 lb.	.06	.08	.10	650
Milk, condensed, 1 can, 14½ oz.	.08	.07	.48	1323
Oleomargarine, 1 lb.	.01	.83	.00	3525
Eggs				
Hens', 1 doz., 1½ lb.	.20	.14	.00	935
Fish				
Bass, 1 lb.	.08	.01	.00	200
Blue, 1 lb.	.10	.01	.00	201
Butter, 1 lb.	.01	.06	.00	460

Foods Primarily of Animal Origin — *Continued*

Food as we buy It	Nourishment in This Food			This Food can yield to the Body in Energy and Heat Units
	Muscle Builder	For Heat and Energy		
	Protein	Fat	Carbohydrates (Starch and Sugar)	
Fish (*con't*)	Pounds	Pounds	Pounds	Calories
Cod, fresh, 1 lb. . . .	.18	.08	.00	165
Cod, salt, 1 lb. . . .	.22	.27	.00	1020
Flounder, 1 lb. . . .	.06	.00	.00	130
Halibut, steak, 1 lb. . .	.15	.04	.00	470
Halibut, smoked, 1 lb. .	.21	.15	.00	1020
Mackerel, 1 lb. . . .	.14	.06	.00	525
Salmon, 1 can, 12 oz. .	.18	.07	.00	638
Shad, 1 lb.	.09	.05	.00	380
Trout, lake, 1 lb. . . .	.09	.05	.00	380
Poultry and Game				
Chickens, broilers, 1 lb.	.13	.01	.00	295
Fowl, 1 lb.	.14	.12	.00	775
Goose, 1 lb.	.13	.30	.00	1500
Turkey, 1 lb.	.16	.18	.00	1475
Lamb				
Chops, 1 lb.	.14	.28	.00	1445
Leg, 1 lb.	.15	.15	.00	900
Mutton				
Kidney, 1 lb.	.17	.03	.00	440
Leg, 1 lb.	.14	.23	.00	1235
Pork				
Bacon, 1 lb.	.13	.36	.00	1740
Chops, loin, 1 lb. . . .	.13	.26	.00	1340
Ham, 1 lb.	.15	.33	.00	1670
Lard, 1 lb.	.00	1.	.00	4226
Sausages				
Bologna, 1 lb.	.18	.20	.00	1170
City, beef and pork, 1 lb.	.19	.24	.00	1380
Frankfurters, 1 lb. . .	.20	.19	.01	1170

Foods Primarily of Animal Origin — *Continued*

Food as we buy It	Nourishment in This Food			This Food can yield to the Body in Energy and Heat Units
	Muscle Builder	For Heat and Energy		
	Protein	Fat	Carbohydrates (Starch and Sugar)	
Shell fish	Pounds	Pounds	Pounds	Calories
Clams, 1 doz., 5 lb.	.10	.00	.05	350
Lobsters, 1, heavy, 3 lb.	.16	.02	.00	400
Oysters, 1 qt., 2 lb.	.12	.02	.06	460
Scallops, 1 qt., 2 lb.	.30	.00	.07	690
Veal				
Breast, 1 lb.	.15	.09	.00	645
Cutlets, 1 lb.	.20	.08	.00	690
Leg, 1 lb.	.18	.06	.00	585
Liver, 1 lb.	.19	.02	.00	575

SUMMARY

Our foods must be both inorganic and organic. Water and mineral matter comprise the inorganic foods. Carbohydrates, fats and oils, and protein, the organic foods.

Protein builds up the tissues and is the only food that can do this.

Carbohydrates and fats (including oils) furnish heat and energy to the body.

A balanced ration is necessary to health.

Food requirement. Under ordinary conditions a man requires $\frac{1}{80}$ of an ounce of protein for each pound he weighs and enough fats and carbohydrates to furnish 2500–3000 Calories daily.

Exercises

1. Name two animal and two vegetable foods rich in (*a*) protein, (*b*) fat or oils, (*c*) carbohydrates.

2. How should a farmer's dietary differ from that of a bookkeeper's? Why?

3. From what foods do we obtain the mineral matter needed for our bodies?

4. Compute the protein you consumed yesterday, and the Calories your food provided. Was your diet well balanced? How could it have been improved?

5. Make out suitable meals for a day for yourself. (Use the one-portion food table.)

6. How does a vegetarian obtain the protein necessary for his body?

7. How would the number of articles of food in a meat and vegetable diet compare with the number in a diet that was purely vegetable?

8. By the use of the family food table determine whether your family is being well and economically nourished.

CHAPTER XXIV

FOOD PRESERVATION

Importance of food preservation. In hundreds of ways science is making our lives easier, and one of these ways is in teaching us how to preserve foods so that we may enjoy meats, fruits, and vegetables, at all times, instead of only in their season.

The table of a laborer is to-day provided the year around with foods that two hundred years ago not even a king could have obtained, and for this we must thank science, which has taught us how to preserve foods from decay. To understand how to prevent the decay of foods we must first understand the causes of their decay.

The fungi. The green coloring matter of plants, called chlorophyll, with the aid of the energy of sunlight, is able to convert the carbon dioxide of the air into the complex compounds found in plants.

One great group of plants, the *fungi,* contains no chlorophyll, and this lack renders them unable to prepare their food from carbon dioxide. They must start from organic matter that has been elaborated by other plants or animals; that is, they are saprophytes. Our common fungi, the mushrooms and toadstools, are examples of such plants. It is to these fungi that the decay and spoiling of foods is almost entirely due.

There are three great divisions of these colorless plants that we must consider, — the molds, the yeasts, and the bacteria. To call these colorless plants is likely to lead to a misunderstanding. They may be highly colored in reds, browns, or grays, but not greens, for the absence of the green coloring matter, chlorophyll, is characteristic of the group.

Molds and their spores. Molds are large enough to be seen by the naked eye. There are many varieties, but as their methods of growth and reproduction are similar, we need not distinguish between them. If you will expose moist pieces of bread, potato, banana, or cheese, to the air, and allow them to stand undisturbed in a warm place for a few days, you will have abundant opportunity to study mold growth (Fig. 99).

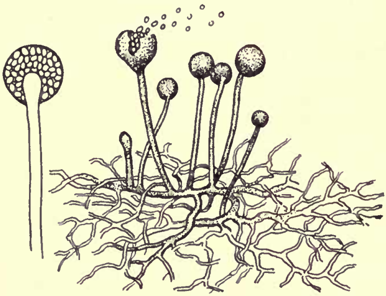


Fig. 99. — Mold, producing spores.

You will notice that first white threads, the *mycelium*, grow down into the host of the mold. These rapidly increase in size and number, and throw out branches until the bread is filled with them. After two or three days, the surface of the mold becomes colored. This color may be a brown, red, or blue, depending on the variety of the mold. This color shows that the mold is fruiting, or producing reproductive bodies called *spores*.

Close examination of the common blue mold, even without the aid of a lens, will show that the mold has thrown up above its surface tiny stalks, each bearing at the top a colored ball. Examination under the microscope shows that these balls contain hundreds of small bodies called *spores*. As the mold matures, these balls break open, and the spores escape to be borne away by the winds, and to cause more trouble to the housewife.

How to avoid molds. The conditions necessary for the germination of the mold spores are moisture, warmth, and food. This gives us the clew to the means of avoiding molds. Jellies are usually covered with a paper or with paraffin. When a spore drops on this covering, it can obtain no food, and so does no harm. If, however, a spore has fallen on the jelly before the cover is put on, it will germinate. Soon the top of the jelly will be covered with a thick coating of mold, and the upper part of the contents spoiled.

Most molds are not harmful, so that the contents of the jar need not be thrown away. If the mold itself, and the soft jelly immediately under it, are removed, the remainder is fit for use. If, after the jelly has been poured into the glasses, and has hardened, the surface is brushed over with brandy, the alcohol will kill any spores present. It may now be covered with the assurance that it will not mold. Covering foods helps to keep them.

Common molds grow best at a temperature of from 70° to 100° F. Keeping food in the refrigerator will then help to prevent the growth of molds. A few varieties, however, grow best at a temperature of 40° F., and these will of course grow freely in an ordinary ice chest.

A temperature of 150° F. will kill any mold that has started to germinate. Some spores will resist this temperature. To make sure, therefore, that canned fruit will not mold, it should be heated to 150° F. Allow it to stand for a day. During this time any spores present will start to grow. Heating a second time will kill these, and the fruit will keep with certainty.

The presence of moisture aids the growth of mold upon almost any organic material. Leather is safe as long as it is dry, but, allowed to become moist, it will mildew, and mildew is one of the forms of mold. Dry food, as crackers, is never troubled with these annoying visitors. The drying of foods is an important method of preserving foods.

The housewife should prevent any molds that may grow in the kitchen from fruiting and scattering spores to make other foods mold. Foods should be kept cool, and where possible in the light, as most molds grow best in the shade.

Most important of all, foods should be kept dry. The importance of this is well seen in apples. Every grocer will tell you that cold storage apples do not keep long after removal from the cold storage warehouse. This is largely because, when they are removed from the cold rooms, they condense a film of water on their skins. Any spores that fall on them will then stick fast. The moisture remains longest where the apples touch each other, and therefore the apples start to decay there. If, after the apples have warmed to the temperature of his store, the grocer would carefully wipe each one, the tendency to spoil would be greatly lessened.

Wrapping fruit in paper tends to preserve it, for not only does the paper keep spores from falling on the fruit, but, being absorbent, keeps the surface of the fruit dry. Knowing the life history of the molds, you can see why one rotten apple will infect a whole barrel. As the mold matures, the spores scatter, and soon all of the fruit is covered with them. Every break in the skin is then an invitation for spores to enter and grow, and they quickly accept the opportunity offered.

Flavors due to molds. Sometimes molds are purposely introduced into food to give it a characteristic flavor. Roquefort cheese when first made is entirely lacking in its well-known taste. To produce this, wires are first thrust into ripe cheese. They are then thrust into the new cheese. In this way the mold from the ripe cheese is introduced into the new cheese, where it grows and produces the green color and pleasant flavor. Brie, Stilton, and Camembert are other cheeses the flavors of which are at least partly due to molds. All of these cheeses should be eaten when just ripe, as the continued growth of the molds finally makes the flavor too strong to be agreeable.

Yeasts and how they grow. Yeasts are microscopic oval plants that reproduce by a process called budding. Examining yeast under the microscope, you can see little more than oval colorless bodies, about $\frac{1}{3000}$ of an inch in diameter (Fig. 100). If you continue your watch, you will see a small knob form on one end of the yeast plant. The cell wall draws in, dividing this knob from the original plant, the two break apart, and we have two plants. In this way yeast plants increase in numbers with great rapidity.

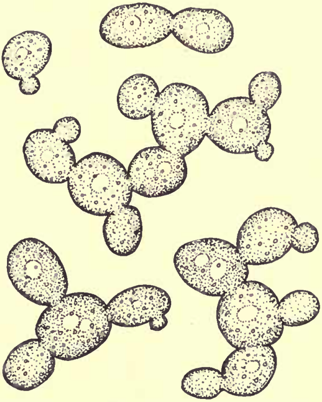


FIG. 100. — Yeast plant.

The enzyme of yeast. The yeast plant secretes a substance known as *zymase*. This is one of a class of bodies known as *enzymes* (also called *ferments*).

When studying oxygen, we found that the decomposition of potassium chlorate took place at a lower temperature if we added manganese dioxide, yet after the oxygen was all given off, the manganese dioxide remained unchanged.

The *enzymes* are compounds that serve a similar purpose, and we may call them by the same name, *catalytic agents.* Their presence enormously increases the speed of certain chemical changes.

There are many of these enzymes or digestive ferments in the body. *Ptyalin* is found in the saliva, and changes starch to sugar. *Pepsin* is found in the gastric juice, and changes proteins to peptones and proteoses. *Trypsin* is found in the pancreatic juice, and has much the same action as pepsin. The particular action that is aided by zymase is the change of sugars into carbon dioxide and alcohol.

$$C_6H_{12}O_6 \rightarrow 2\ C_2H_5OH + 2\ CO_2$$

Yeast in bread making. The use of yeast in making bread light has been known from antiquity. Leavened bread was bread made light by adding leaven to the dough. This leaven was dough saved from a previous baking, and contained yeast.

Yeast spores are found floating in the air everywhere, and if dough is exposed to the air, some of these "wild yeasts" are sure to fall into it. These wild yeasts are not always desirable additions, for some of them cause the bread to taste sour.

From these wild yeasts the cultivated yeast of to-day has been obtained. The yeast plants are placed in large vats, mashed potato and hops added, and the whole kept at a temperature of 80° F. Yeast plants multiply rapidly, the contents of the vat becoming thick and frothy. The yeast is then thrown on a cheesecloth, to remove some of the excess water. When well drained, starch is mixed with it to make it firmer, and it is cut into cakes and sold as compressed yeast.

If yeast is gently dried, it passes into a dormant state, in which state it is sold as dry yeast cakes. These are used where a supply of yeast must be kept for some time, as compressed yeast will not keep long.

The carbon dioxide produced by the action of yeast in dough makes the bread porous, and more easily digested, owing to the larger surface exposed to the action of the digestive juice. Biscuits raised with baking powder, and aërated bread, contain the same ingredients as ordinary bread, but have a very different flavor.

Fermented drinks. All fermented drinks, from homemade root beer to champagne, are made by the action of yeast on sugar. In root beer a small quantity of a highly flavored extract is mixed with sugar, yeast, and water, and bottled. Fermentation occurs, and the carbon dioxide produced makes the liquid bubble and froth when the bottle is opened.

It must be remembered in using these homemade beverages, that alcohol is produced in addition to the carbon dioxide, and that, if the fermentation continues for several days, an amount of alcohol is produced as great as that contained in a light beer.

Kumiss is made by adding yeast and sugar to milk. The mixture is allowed to ferment for a day, and then bottled. When cool it is ready for use. It is claimed that it is more digestible than raw milk, and is used by invalids.

Manufacture of grain alcohol. Ethyl alcohol, or grain alcohol, C_2H_5OH, is made by converting the starch of grain into sugar, fermenting this, and distilling the alcohol. Grain, as barley, is spread out on floors to a depth of six inches. It is kept moist and warm. The grain starts to germinate, when the diastase present changes the starch of the grain to sugar.

As soon as the maximum amount of sugar has been obtained, the grain is ground, and yeast and water added. Fermentation takes place, and alcohol is produced. The mixture is distilled, and as alcohol boils at a lower temperature than water, the distillate contains all of the alcohol, mixed with a good deal of water. The dilute alcohol is then again distilled to concentrate it. In this way the 95% alcohol of the druggist is prepared.

Properties of alcohol. Alcohol is a colorless, volatile liquid, lighter than water, and miscible with it. It burns with a very hot, almost colorless, blue flame. It is the intoxicating constituent of all fermented drinks. It is widely used as a fuel, and as a solvent. Resins and oils dissolve in it freely. Shellac varnish is a solution of shellac in alcohol. The tinctures of the druggist, as tincture of iodine, are all solutions in alcohol.

Denatured alcohol. The government levies a heavy tax on alcohol that can be used for drinking purposes. Formerly this tax was a burden on the manufacturer of chemicals, who uses large quantities of alcohol as a solvent. To meet industrial needs, the government allows what is called *denatured alcohol* to be sold tax free. This is alcohol to which some substance has been added that renders it unfit for drinking, but does not interfere with its use as a fuel or solvent. Denatured alcohol can be sold at about one fifth the price of pure alcohol, as it pays no tax.

Wood alcohol. When wood is distilled, one of the products obtained is wood alcohol, or methyl alcohol, CH_3OH. This is a light, colorless liquid, having a disagreeable odor. It is used, like grain alcohol, as a fuel and a solvent. It is a poison, and must not be used in beverages.

Ether. By distilling a mixture of strong sulphuric acid and alcohol, ether is obtained.

$$2\ C_2H_5OH + H_2SO_4 \rightarrow (C_2H_5)_2O + H_2SO_4 + H_2O$$

Ether is a very volatile, colorless liquid, having a characteristic odor. Its vapor is very heavy and inflammable. Ether is used by surgeons as an anæsthetic.

Manufacture of vinegar. When sweet cider is exposed to the air, it first turns hard; that is, yeast converts the sugar into alcohol. If this hard cider is allowed to stand, it is converted into cider vinegar. This is due to the action of mother-of-vinegar. An examination of mother-of-vinegar under the microscope shows that it is composed of myriads of bacteria. Vinegar, then, which is a weak solution of acetic acid, is one of the products of bacterial action, and we should expect that any dilute solution of alcohol would be acted upon in the same way.

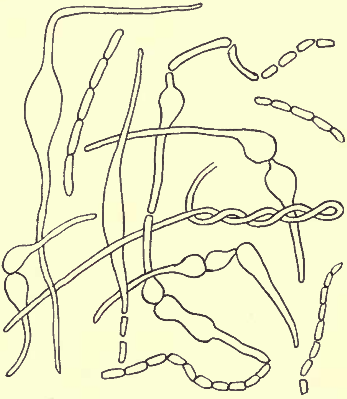


FIG. 101. — Bacteria in mother-of-vinegar.

Much of the vinegar used in pickling is prepared from alcohol. It is the ordinary "malt vinegar" or "white wine vinegar" of trade. A large cylinder is filled with wood shavings, placed on trays so that they cannot pack too closely together. These shavings are first soaked in vinegar, to impregnate them with mother-of-vinegar. Provision must be made by means of holes for the passage of air through the cylinder, for the change of alcohol to acetic acid is really an oxidation of the alcohol (Fig. 102).

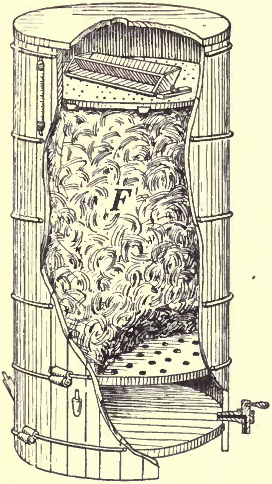


Fig. 102. — Vinegar making. Quick process.

Dilute alcohol is then allowed to trickle slowly through the cylinder. As it passes over the shavings, which are covered with the necessary bacteria, the alcohol is converted into acetic acid. If the conversion is not complete by one passage through, the liquid that runs from the bottom of the cylinder is returned to the top and run through again.

The product is satisfactory for cooking and pickling processes, but lacks the flavor of cider vinegar. It has sometimes been a temptation that the manufacturer could not resist to add a little caramel to the product. This gives it the color of cider vinegar, and it is sold under that name.

Bacteria and how they reproduce. Bacteria are still smaller than either molds or yeasts. They are so small indeed that the highest powers of the microscope are needed to see them, and even these high powers do little more than show us moving dots. One of the largest of these bacteria is about $\frac{1}{10000}$ of an inch long, while many are not more than $\frac{1}{50000}$ of an inch in diameter.

It is only within the past few decades that we have known anything about these minute forms of life, for they are invisible to the eye, and their study even under the microscope tells us little about them. They are very simple in structure, and it is quite impossible in many cases to identify the variety by looking at them. We must test them to see

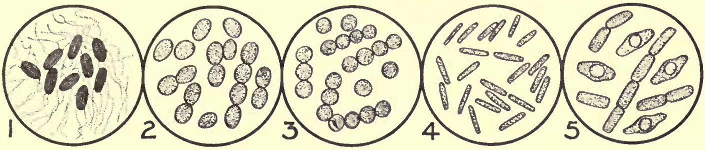


Fig. 103. — Bacteria. (1), Typhoid fever. (2–5), Forms of bacteria found in milk.

what they can do. In form, they are usually spherical, rod-shaped, or spiral.

Bacteria multiply by simple division. A rod-shaped individual will lengthen, and then divide in two. This process is named *fission*. Reproduction is very rapid, some bacteria dividing every half hour. This accounts for their immense numbers, for if every individual divides every half hour, at the end of six hours there would be 4096, and at the end of 24 hours about 17,000,000 bacteria all descended from one parent.

These organisms are plants, but have the power to move. In popular usage bacteria are called *germs*, and all kinds of malign influences are attributed to them. This idea is not quite right. It is true that certain varieties, called the *pathogenic bacteria*, cause certain diseases. Thus, consumption, typhoid fever, and diphtheria are caused by the presence in the body of certain bacteria. The pathogenic bacteria are, however, in a small minority, the nonpathogenic forms far outnumbering them.

These nonpathogenic forms are of great importance to us. They destroy the organic refuse of the world, and return its elements to the soil in such a form that plants can use them. They are of great importance to the farmer. One form in particular fixes the nitrogen of the air, that is, changes the free nitrogen of the air which plants cannot assimilate into a nitrate which they can use.

These nitrifying bacteria are often found associated with the *legumes,* such as peas, beans, lentils, clover, and alfalfa. They grow in nodules on the roots of these plants and make the soil fertile. In starting a new field of alfalfa, the ground should always be inoculated with these nitrifying bacteria before planting the alfalfa seed. This is done by sprinkling dirt from an old alfalfa field over the ground of the new field. Bacteria are also useful to the dairyman, for the flavor of butter and cheese depends on their action, as well as on that of the molds.

Some pathogenic bacteria cause diseases to which only certain forms of animal life are subject. Advantage is sometimes taken of this to exterminate rats and mice. Food is spread around that contains these bacteria. The mice and rats eat it, and die of the disease that it produces. Other animals can eat the food with impunity, as they are not subject to the disease caused by these bacteria.

How to control bacteria. To the housewife, however, bacteria are always a trouble, for it is their presence and activity that make the preservation of food difficult. We must then study the conditions that favor their growth, so that we may know how to control them. Most of the common forms grow best at the ordinary temperature, or a temperature slightly higher. A few, however, grow best at a temperature only slightly above freezing. Keeping food cold will then help to preserve it.

Certain kinds of bacteria multiply by spores. These are formed inside of the individual bacteria, break through the cell wall, and float around until they find a suitable growing place. These forms are very difficult to get rid of, for their spores are so resistant that even boiling them for a short time will not kill them.

Boiling will kill practically all bacteria. Their spores, however, will resist boiling for a short time, so that to be sure that food is *sterile*, that is, contains no bacteria, we must either boil it for an hour, or boil it for a short time, allow it to cool and stand, so that any spores present will develop, and then boil again. Merely warming the food will not answer, for this only causes the bacteria to grow the faster.

Light is a great enemy of bacteria. They grow best in dark, damp, dirty corners. To keep our houses as free as possible from these enemies, we must have plenty of sunlight streaming into every room. The sunlight may fade the carpet, but this is of no consequence, compared to the preservation of our health. Light is especially fatal to the pathogenic forms. In sickrooms, then, the common practice of keeping the room dark is not to be recommended. A light room will not only make the patient more cheerful, but will assist in controlling the disease by killing the germs that may cause it in others.

Like molds, bacteria require water for their growth. There is this difference, however: molds grow best on acid foods, while bacteria grow best on slightly alkaline materials. Fruits therefore decay by the action of molds, while meats spoil by the action of bacteria. This rule is not of universal application, but is generally true.

Where bacteria are found. Bacteria are found everywhere. An examination of the saliva, of the dust of a room, or of a bank note, will show their presence in large numbers. Small as they are, they are somewhat heavier than the air, and therefore settle to the floor, where they can easily be found in quantities.

In cleaning a room the common practice of sweeping, and then dusting with a feather duster, simply stirs up the bacteria, and they find new and less desirable growing places. It is better to use an oiled cloth and wipe the dust away.

Many diseases are spread by a disregard of these facts. The saliva of a consumptive contains the bacteria that cause the disease. If, then, the consumptive spits in the street, and the saliva dries, many of these bacteria will be set free to float in the air, to be drawn into the lungs of healthy persons, and so expose them to the danger of infection. Too great care cannot be exercised in keeping these pathogenic forms under control.

Action of bacteria on proteins. Proteins form the main food of bacteria. Their action causes its putrefaction and makes the food unfit for use. Even food that shows no outward sign of their action may be dangerous, for, in addition to the products arising from the decay of the food itself, bacteria secrete some poisonous compounds. These are called *ptomaines*. Fish, shellfish, and ice cream are very likely to develop ptomaines, especially in the summer. They often cause severe and sometimes fatal poisoning.

Pasteurization of milk. Different varieties of bacteria produce different results. Thus, the mother-of-vinegar bacteria change alcohol to acetic acid. One other variety is troublesome to the housewife, the variety that changes milk sugar to lactic acid, and thus causes milk to sour. This variety does not form spores, and is killed by exposure to a temperature of 155° F.

To keep milk sweet, then, it should be heated to 155° F. This process is called *pasteurization*. Milk heated to this temperature does not have the taste of boiled milk, nor are the albuminoids coagulated, as they are in boiled milk. To pasteurize milk it should be placed in bottles, and these placed in a double boiler and heated to from 155°–170° for twenty minutes (Fig. 104). The milk should then be cooled rapidly, and kept cold. If pasteurized milk is kept at ordinary temperatures, it will spoil more quickly than raw milk. Although pasteurized milk does not sour, it is subject to putrefaction and should not be kept over 24 hours.

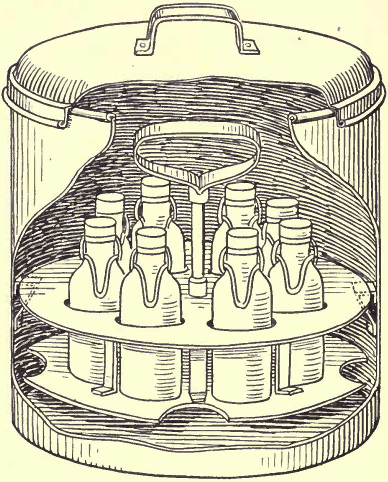


FIG. 104. — Pasteurizing apparatus.

Preservation by cooling and drying. One of the best means of preserving food is to keep it cold. The use of refrigerators, and of cold storage, is universal for this purpose. Even cold storage will not, however, keep food indefinitely, and a limit should be placed on the length of time food may remain in cold storage. As you learned in the previous paragraphs, cold storage is efficacious because most molds and bacteria do not grow in the cold.

Keeping food dry also tends to preserve it. Dried fruits, fish, and meats are all examples of the use of this method of preserving food. Its usefulness is due to the fact that bacteria and molds require moisture for their growth.

Preservation by canning. Boiling will preserve foods, for this kills the bacteria, and if the food is protected from fresh infection, it will keep a long time. The canned vegetables that form such a large proportion of the stock of the grocer are examples of this method of preservation.

The vegetables are washed, cut up if necessary, and placed in tin cans. The cans are then covered and heated, often under pressure to raise the boiling temperature, until the vegetable is cooked. The cans are then sealed. The difficulty in the process comes in being sure that *every* form of microörganism is killed by the heating. If a single spore survives, the contents of the can will spoil, and, as some of the bacteria spores are very resistant, it is difficult to be sure that they are all killed. Properly canned foods keep indefinitely, and the process has been a boon to humanity.

The canning of fruit at home depends on the same principles. If the fruit jars, the rubbers, and the tops are sterile, and the fruit is heated until it too is sterile, the fruit will keep. If, however, the rings are washed in cold water, or the hot fruit allowed to stand open in the jars for a while, it will probably not keep. Can you now explain why preserves sometimes "work," and why, when they are opened, bubbles of gas are seen and the preserves have an alcoholic smell?

Chemical preservatives. Foods are sometimes kept by the use of chemical preservatives. These must have two properties. They must have antiseptic powers, and they must be harmless to man. There are many substances such as mercuric chloride, corrosive sublimate, that will prevent the growth of the microörganisms that cause food to spoil. Corrosive sublimate is, however, a violent poison, and so cannot be used. Other substances as sugar, salt, spices, vinegar, and the smoke from smoldering wood, have a preservative action, and are regarded as harmless. Still another class of preservatives, including borax, boracic acid, benzoic acid and benzoates, salicylic acid and salicylates, and sulphites, is used.

Some of these, used in small quantities, are probably harmless. Their use, however, makes it possible for the manufacturer to use material that is unfit for human consumption. If tomato catsup can be made without the use of any preservative, why should we buy a product that contains a preservative? It may be harmless, and the material used by the maker may be of a high grade, but so long as its use makes it possible to use inferior materials, and so long as it is not certain that preservatives are harmless, why buy the doubtful products?

There is another objection to their use. Even the manufacturers who use these preservatives claim only that they are harmless in small quantities. Now if, as sometimes happens, much of the food served at a meal is canned, and if each of the articles contains only a little preservative, the total amount that we consume may be considerable. It is therefore a wise rule to avoid their use altogether. *Never buy food preserved with sodium benzoate or other preservatives, even if their use is allowed by law.*

One notable example of the use of such harmful preservatives is the use of formaldehyde in milk. This will preserve the milk, that is, the milk will not turn sour, but it will rot. Its use is illegal, but some cheap milks contain it. If milk will not sour in the ordinary length of time, you may be sure that a preservative has been added. Never buy or use such milk.

The same thing is true of all foods. Whenever you find a brand of catsup that will not spoil, or sausages that do not go bad, do not use them. They must contain preservatives, and it is best to avoid all such food products.

Antiseptics. The same agents used in the preservation of foods may also be used as disinfectants, and here we must be sure that we understand the meaning of certain terms. A *deodorizer* is a substance that destroys a noxious odor or covers it up with a stronger one. Many deodorizers are of no value, indeed are harmful, for their use leads to the belief that the cause of the odor has been removed, which is not the case. A liberal use of cologne may cover up the odor of perspiration, but it will not obviate the necessity for bathing.

An *antiseptic* is a material that checks the growth of bacteria. It may not, however, kill all of those present. A *germicide* is a material that kills all bacteria. Antiseptics, when used in quantities, are generally germicides.

To *disinfect* is to remove the cause of trouble. If chlorine oxidizes decaying matter, it is a disinfectant. If corrosive sublimate is used in a strength that will kill all bacteria and prevent putrefaction, it is a disinfectant. We have already seen that both heat and sunlight will kill bacteria. We may use both of these as disinfectants.

The number of chemical disinfectants is great, but there are five that can so easily be used at home that you ought to know something of them. One of the most important is mercuric chloride, called corrosive sublimate, $HgCl_2$. Used at the strength of 1 to 1000 of water, it is quickly fatal to all microörganisms. It is intensely poisonous, and this limits its use to places where there is no danger of its being swallowed.

Bleaching powder, or chloride of lime, is extensively used. Its value depends upon the liberation of chlorine. When the damp material ceases to have a strong odor, it is exhausted, and should be replaced by fresh.

Carbolic acid, or phenol, is an excellent disinfectant. It is not as efficient as either corrosive sublimate or chloride of lime, but is used more than either. It is an intense poison, and must be handled with care.

Formaldehyde. Formaldehyde, or formalin, is one of the best disinfectants for home use. Formaldehyde is a gas with a disagreeable odor. It is soluble in water, and a 40% solution of it is sold as formalin. It has no effect on fabrics, and so clothes that need to be disinfected may be soaked in it.

One of the difficulties of effectively disinfecting a room is to make sure that all parts of the room come into contact with the disinfectant. Formaldehyde is a gas, and, if it is liberated in a room in quantity, every part of the room will be disinfected.

Formaldehyde polymerizes, that is, several molecules combine together to form a new molecule. This polymer is called paraformaldehyde. It is a white solid, which on heating changes back into the gas formaldehyde. The solid is heated over a lamp and is thus made to give off the disinfecting gas. This is an easy and efficacious way of disinfecting a room.

Sulphur dioxide. Sulphur, when burned, forms the gas sulphur dioxide, SO_2. This gas is easily oxidized, forming sulphur trioxide, SO_3. It is therefore a good reducing agent. For this reason it is used as a bleaching agent, especially for straw, and dried fruits. A mass of sulphur, burned in a closed room, is a fair disinfectant, but not so good as formaldehyde. It has the advantage that sulphur can be bought anywhere, and its use requires no especial apparatus.

The gas sulphur dioxide is easily compressed to a liquid and is sold in the liquid state in tin cans having an opening closed by a lead pipe. To use these cans to disinfect a room, it is only necessary to close the room as tightly as possible, sealing all cracks by pasting paper over them, then cut the lead pipe. This allows the sulphur dioxide to flow out. Place the can on a tin dish, go out, and close the door. The liquid changes to a gas, which, going everywhere in the room, disinfects it.

SUMMARY

Molds cause foods to spoil. They propagate by spores. They require moisture, warmth, and food, before they can grow. Heat kills them. Certain foods, as cheese, owe their flavor to them.

Yeast contains zymase, which changes sugar to alcohol and carbon dioxide.

An enzyme is an organic catalytic agent which hastens certain specific chemical changes.

Ethyl alcohol, or grain alcohol, is prepared by the action of yeast on sugar. The product is distilled to purify it. It is used in beverages, as a fuel, and as a solvent.

Denatured alcohol is alcohol made nonpotable by the addition of wood alcohol, etc.

Ether is ethyl oxide. It is made by taking one molecule of water from two molecules of alcohol. It is used as an anæsthetic.

Vinegar is made by the action of mother-of-vinegar on alcohol. It is dilute acetic acid.

Bacteria are microörganisms. They multiply with enormous rapidity by *fission*. They cause foods, especially protein, to putrefy, and are useful as they destroy organic waste.

Pathogenic bacteria cause disease. Boiling kills them.

Ptomaines are poisonous bodies produced by certain bacteria.

Pasteurization is heating milk to 155° F. This kills the bacteria that produce lactic acid and sour the milk.

Keeping food dry and cold preserves it, because this prevents the growth of bacteria and molds.

Harmless preservatives: sugar, salt, spices, vinegar, and the smoke from smoldering wood will preserve food. These preservatives are harmless.

Chemical preservatives to be avoided: borax and boracic acid, benzoates and benzoic acid, salicylates and salicylic acid, and sulphites, will preserve foods. They are all probably somewhat harmful. Formaldehyde is a poison and should never be used as a food preservative.

A deodorizer covers up a bad odor with a stronger one. It is of no value.

An antiseptic is a material that checks the growth of bacteria.

A germicide is a material that kills all bacteria.

Household disinfectants: mercuric chloride, chloride of lime, carbolic acid, formaldehyde, and sulphur dioxide.

Exercises

1. Why is it important to prevent molds from forming spores?
2. Why do preserves sometimes "work"?
3. Why do not preserves in strong sirup spoil, even if left open?
4. Why do canned soups sometimes spoil?
5. What would you conclude if bottled milk kept sweet for four days?
6. Why does sweet cider become hard?
7. Name two enzymes and explain what each does.
8. If vinegar is only dilute acetic acid, why not use dilute sulphuric acid instead? It would be cheaper.
9. How can druggists keep "fresh fruit" for soda fountain use on their counters for days without its spoiling?

CHAPTER XXV

SILICON, SILICA, AND SILICATES

Silicon. It is remarkable that the second most abundant element is not known, even by name, to the majority of people. Silicon constitutes 28% of the earth's crust, and yet you probably never saw even one specimen of the pure element. This is because it always exists naturally in compounds, either as the dioxide, or as a silicate. It resembles carbon in its chemical properties.

It is reduced from its compounds with great difficulty. But with the aid of the electric furnace it is now made in large quantities. The element is a brown, crystalline body, which burns readily in the air. The element itself is of no importance in the home, but its compounds are in use by every one.

Silica and its varieties. Silicon dioxide, SiO_2, known also as silica, occurs distributed everywhere in nature. In the form of quartz it constitutes the bulk of sand, and of many soils. Many plants contain it; the coating of rattan, for example, is almost pure silica.

When pure, it crystallizes in six-sided prisms, topped by six-sided pyramids (Fig. 105). These clear, colorless crystals are called rock crystals and, when cut, are rhinestones. Many of the stones sold under such names as "Congo diamonds" are cut from quartz.

Quartz is harder than glass, and has only a small coefficient of expansion. That is, it expands very little on heating. A dish made of quartz can therefore be heated red-hot and then plunged into ice water without breaking. This, together with the fact that quartz is infusible at a white heat and is unaffected by most chemical reagents, makes it valuable in the laboratory for dishes and crucibles. These were formerly so expensive that chemists could not afford to use them, but they are now made cheaply by the use of the electric furnace.

Fig. 105. — Quartz crystals.

The colors in quartz are due to small amounts of impurities. A little manganese colors it violet, and makes the semiprecious stone, the amethyst. Other impurities produce smoky quartz, or the cairngorm stone of Scotland, rose quartz, chalcedony, jasper, carnelian, and the so-called yellow topaz.

Opal is quartz containing some water. Its play of color is due to fine cracks in the stone, which cause an interference of light. These colors are produced in the same way as are the colors in a soap bubble.

Many of these stones when cut are of a beautiful color. They all have, however, the disadvantage of being soft, compared to the precious stones, and so do not hold their polish when subjected to hard wear.

Sand is broken-up quartz, and, when cemented together, makes sandstone. Flint, too, is a variety of quartz. Quartz sand is used in making glass, porcelain, and as a building material in mixing mortar and cement.

Silica as an abrasive. Quartz is largely used as an abrasive and polishing material. Sandpaper consists of sharp sand glued to paper. The shells of certain small water plants, the diatoms, are largely silica. When they die, their soft parts decay, but the shell being silica remains, falls through the water, and forms a layer of mud on the bottom of the water. The shell remains of these diatoms are called diatomaceous earth, or infusorial earth (Fig. 106). When freed from impurities, it is sold as a polishing powder.

Fig. 106. — Diatoms. (× 50.)

Examine some "electro-silicon" under the microscope, and you will see the curious structure of the silica valves of these diatoms.

Agate. Quartz is insoluble in pure water, but in the presence of an alkali it will slowly dissolve. Since many natural waters, especially in the western part of our country, are alkaline, quartz dissolves in them. This dissolved quartz may be precipitated by an acid. It is in this way that agates are formed. Water containing silica runs into a hole in the rock, the quartz is deposited, and forms a layer. Sometimes the water is clear and leaves a clear layer.

Sometimes it is muddy and leaves a brown or black ring. In this way the banded structure of the agate is produced. The brilliantly colored agates are all colored artificially.

Common silicates. Silicon forms a large number of acids known as silicic acids. Their salts, called silicates, make up the bulk of the rocks of the earth. Feldspar, slate, soapstone, serpentine, asbestos, and talc are all silicates. We may easily prepare sodium silicate in the laboratory, by fusing together sand and sodium carbonate, and dissolving the resulting mass. Sodium silicate is called water glass, or soluble glass, and is used as a cement for asbestos and glass, in preserving eggs, and by calico printers and soap makers.

Making glass. Among the artificial silicates the most important is glass. By melting together in a large clay pot, sand, calcium carbonate (marble), and sodium sulphate or carbonate, a transparent viscous mixture of sodium and calcium silicates is obtained. This is glass.

If the materials used are perfectly pure, a colorless glass results. The presence of iron as an impurity gives us green bottle glass. By purposely adding a coloring agent, we can produce glass of any color. Cobalt gives a blue, manganese, a violet, and chromium, a green color to glass.

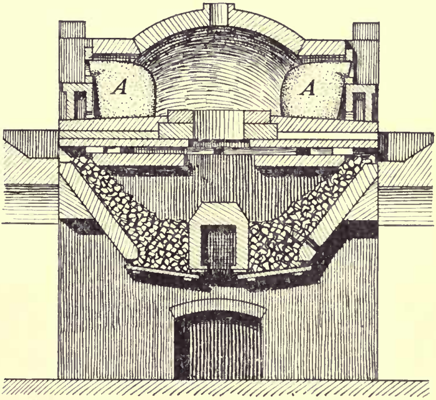
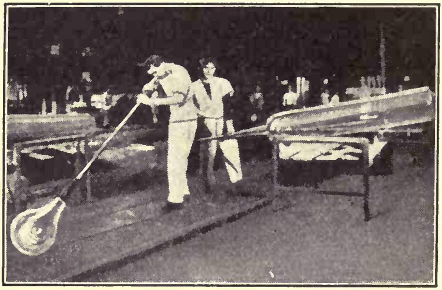


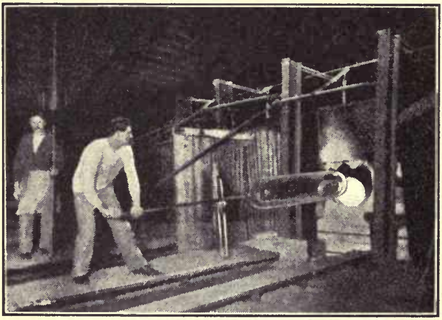
FIG. 107. — Glass furnace. The glass is melted in clay pots (*AA*) which are surrounded by hot gases.

Many glass utensils are blown. The workman pushes the end of an iron tube (the blowpipe) into the mass of molten glass, and collects a ball of the thick material on the end of his blowpipe. By blowing through the blowpipe, this ball is changed into a bubble of glass, and, as it is soft, the workman can form it into any required shape, as a bottle or tumbler.

Window glass is made by blowing a large bubble and then swinging it until it stretches out into a cylinder (Fig. 108). The top and bottom of this cylinder are then cut off, and the cylinder cracked down one side. It is then placed on a hot iron plate, the cracked side uppermost. The heat softens the glass, and the cylinder slowly opens, forming a sheet. This is flattened as much as possible by pressing it against the iron plate with an iron instrument resembling a spade. Naturally the sheet of glass is never quite flat, hence the unevenness of our window glass.



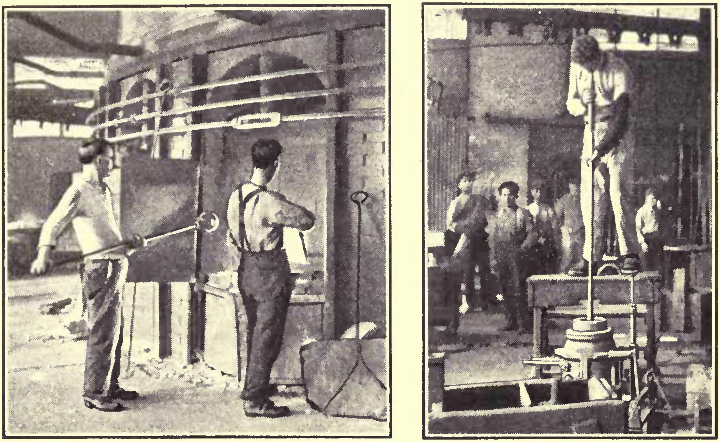
(*a*) Blowing the cylinder.



(*b*) Reheating.

Fig. 108. — Window glass making.

Plate glass. If we wish perfectly flat glass, we must use plate glass. This is made by pouring molten glass upon an iron table having a ledge all around it to keep the glass from running off. A heavy roller is then passed over it to flatten it. This gives a thick sheet of glass. It is then ground and polished on both sides. As much of the glass is wasted in the grinding and polishing, and as the cost of these two operations is high, plate glass is much more costly than ordinary glass.



(*a*) Getting a lump of glass on the blowpipe. (*b*) Blowing a bottle in a mold.

FIG. 109. — Bottle making.

Molding glass bottles. Many bottles are molded, instead of being blown into shape. The blowpipe, having a lump of glass on its end, is inserted into a mold. The workman blows through the pipe, thus forcing the glass to take the shape of the mold (Fig. 109). The mold is then opened, and the completed bottle taken out. Usually a ridge left by the place where the mold opens can be seen on such bottles.

Any variation in the amount of glass used will alter the thickness of the walls of the bottle, and this will alter its capacity. To be sure that the bottle will hold the required amount, it is weighed. An automatic weighing machine is used that will reject the bottle unless its weight falls between certain fixed limits. In this way the bottles that are too thin or too thick are detected and remelted.

Making glass tubing. Glass tubing is made by blowing a small bubble of air into a ball of glass on the end of a blowpipe. A blowpipe is then stuck in the side of the glass ball opposite to the first, and the two workmen run in opposite directions. This stretches the ball out into a tube. By varying the size of the air bubble in the original ball of glass, and the speed at which the men run, large or small tubes, with thick or thin walls, can be made.

Chemical glass. Many glasses contain silicates of other metals than sodium and calcium. For chemical glassware a potassium calcium glass, called Bohemian glass, is used, as it is harder and more infusible than the sodium calcium glass. For the glass used in photographic lenses, the most diverse mixtures of silicates are used.

Cut glass. Cut glass is made from a lead glass, that is, glass containing lead silicate. It is heavy, and takes a beautiful luster on being polished. To make a cut glass bowl, the heavy bowl shape is first molded. The design is then drawn on this with chalk and then cut on sandstone wheels. This leaves the glass rough and like ground glass. It is then polished, using rouge or some similar polishing powder. It is this cutting and polishing that make cut glass expensive.

Ground glass. Ground glass is made by the use of a sand blast. Sand is blown by compressed air against the surface of ordinary window glass. Each grain of sand as it strikes the glass makes a tiny scratch, and these scratches crossing and recrossing each other make the glass translucent. Fine ground glass is also made by pouring dilute hydrofluoric acid on a glass plate.

Graduating glassware. Glass resists the action of most chemicals, but hydrofluoric acid, HF, dissolves it. Advantage is taken of this in marking thermometers and graduates. The stem of the thermometer is first covered with wax. This is scratched through with a needle where the graduations are to come, and hydrofluoric acid poured on. This dissolves the glass wherever it has been exposed, and has no action on the wax. In this way lines or letters are placed on glass.

Formation of clay. So many minerals are silicates that a mere list of their names would fill several pages of this book. Meerschaum is a silicate of magnesium; garnet and hornblende are both silicates of calcium and magnesium; mica is a silicate of potassium and aluminium; feldspar is a silicate of aluminium and sodium or potassium; and asbestos is a silicate of magnesium. These are only a few examples of the many important common silicates.

The familiar rock, granite, used in building and for monuments, is composed of a mixture of quartz, feldspar, and usually mica or hornblende. Long-continued action of air and water eventually breaks up the granite. Quartz is set free as sand; mica, being light, is washed away; and feldspar remains. This feldspar disintegrates, the sodium or potassium silicates wash out, and aluminium silicate is left. This is called clay. When free from all impurities, it is a white, earthy mass, called kaolin, $Al_2Si_2O_7 \cdot 2\,H_2O$.

Brick, tile, and stoneware. Clay when heated becomes a hard mass. Bricks and tile are made from it by this process. When clay is mixed with water, it becomes plastic. This plastic clay is molded into the shape of the brick and set aside to dry. When dry it is "fired," that is, heated to a red heat in a kiln. Since the clay used contains fusible impurities, these melt, and bind the mass together into a hard brick. If iron is present in the clay, as is usually the case, the brick turns red. If the clay is pure, a white brick results.

Stoneware and flower pots are made by the same method as brick. To glaze them salt is thrown into the kiln while the stoneware is being fired. This volatilizes, and attacks the surface of the pot, forming a fusible sodium silicate. This melts and spreads over the surface of the ware, thus forming a smooth, impervious layer called a glaze. Such stoneware as this is cheap, but the glaze is often imperfect, and the ware itself is porous. It is not suitable for plates, cups, saucers, and similar ware.

Fig. 110. — Inside a pottery kiln.

Porcelain. The best quality of ware made from clay is porcelain. To make this, pure clay is mixed with fine sand and feldspar, and the whole is ground to a fine powder, which is mixed with enough water to make a plastic mass. A lump of this is placed on a potter's wheel, which is a horizontal round board that can be rotated. The potter rotates the wheel, and with his fingers shapes the lump of clay into a vase, plate, cup, or any other object desired. The article is then air dried and fired. The feldspar contained in the mixture melts, and changes the whole mass into translucent, white porcelain.

The glaze used with porcelain is a finely ground mixture of much the same composition as porcelain itself, but more fusible. This is mixed with water, and the fired porcelain dipped into it. A thin coat of glaze is left all over the article. The procelain is then fired again, the glaze melts, and forms the smooth outside surface that we see on our table porcelain.

When a considerable number of articles are wanted that are to be duplicates, instead of " throwing " the clay on the potter's wheel, it is molded or cast. Plates, cups, and saucers are commonly made in this way.

Let us suppose that we wish to cast a cup. A plaster of Paris mold, having the shape on the inside that the cup is to have, is made in two pieces. The procelain mixture is stirred with water until a thick cream called *slip* is produced. The halves of the mold are then tied together and the slip poured in. The plaster, being porous, absorbs water, and a thin layer of the porcelain mixture sticks to the mold. When this layer is thick enough, the excess slip is poured out, the halves of the mold separated, and the cup taken out. A piece of the plastic mixture is rolled out, and from this the handle of the cup is formed. This is pressed on the cup and sticks fast. It is air-dried and fired.

Porcelain may be decorated by painting any design on its surface, using a colored fusible silicate as paint. When the procelain is fired, this paint melts into the glaze, and becomes a part of it.

Cement and reënforced concrete. Cement is made by heating together limestone, clay, and sand in the proper proportions, and grinding the resulting clinker to a dust. When this is mixed with water, a mass is produced that becomes hard, even under water. Sometimes a rock is found in nature containing all three necessary ingredients in the proper proportion. Such a rock occurs near Naples, and was used by the Romans to produce the cement from which they made their famous aqueducts.

When cement is mixed with sand and crushed stone, and then water is added, the mass hardens to an artificial stone called *concrete*. This is extensively used for sidewalks, building foundations, and floors. Concrete buildings usually contain iron rods imbedded in the concrete. This makes what is known as *reënforced concrete*. Such buildings are almost everlasting.

Blast furnace slag, when mixed with limestone, heated, and ground, makes an excellent cement. This is another good illustration of the way that chemists are utilizing materials which were formerly waste products.

SUMMARY

Silicon is the second most abundant element.

Silicon dioxide, **silica**, or **quartz** is a common mineral.

Rock crystal, **amethyst**, **agate**, **jasper**, **carnelian**, **onyx**, are all varieties of quartz.

Sand is silica.

Sandstone is sand cemented together.

Opal is quartz containing water. Its fire is due to cracks in the stone.

Sand paper is sharp sand glued on to paper.

Diatomaceous earth is the silica valves of diatoms. It is used as a polishing powder.

Glass is a mixture of silicates. Window glass is sodium calcium silicate. Cut glass contains lead silicate. Plate glass is ground flat and then polished.

Hydrofluoric acid will attack glass.

Pottery is made by burning clay.

Cement is clay, limestone, and silica ground together and heated. Mixed with sand and crushed stone it is *concrete.*

Exercises

1. How could you distinguish stoneware from porcelain?
2. How is cut glass distinguished from its imitations?
3. What advantages has a concrete house?
4. Why are flower pots red, while clay is white?
5. How can letters painted on glass bottles be made permanent?

CHAPTER XXVI

TEXTILES

Fibers used in cloth. Varied as are the fabrics and weaves used in dress goods, the fibers used to produce them are few in number. Silk, wool, cotton, and linen are the only four of much commercial importance. Many others are used in small amounts, as hemp, ramie, asbestos, and jute.

Characteristics fibers should have. To make it profitable to spin fibers, they must possess a number of characteristics which adapt them to use in thread or fabric. Cloth cannot be woven until the thread is spun, and to spin suitable thread, fibers having a considerable tensile strength are necessary. Fibers which are able to resist a pull give good wearing qualities to the cloth.

The fibers, too, must be long, as otherwise the thread will not hold together. The separate fibers must have some ability to stick together, as otherwise the thread is weak. They must be pliable, so that they may be easily bent in weaving. They must be porous, or they will not take a dye, nor can they be easily bleached. They must be thin, as otherwise we cannot weave a fine cloth. They must be of uniform size, as otherwise the cloth will not be of uniform thickness. Lastly, to make weaving a commercial success the fiber must be available in large quantities, must be cheap, and must wear well.

If you will consider the above necessary qualities that a fiber must have before it is suitable for commercial

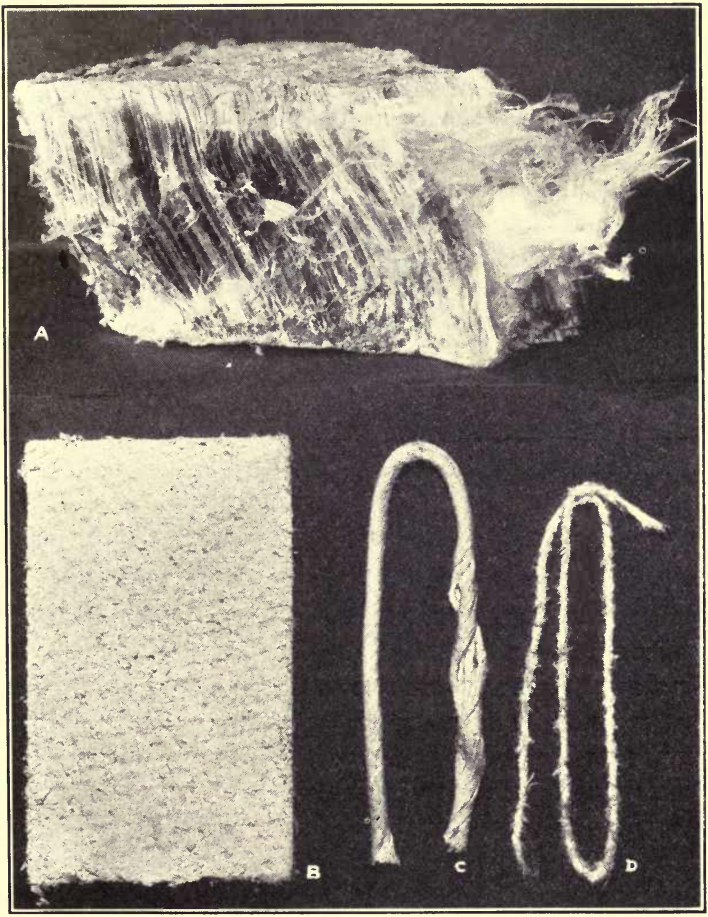


Fig. 111. — Asbestos; *A*, native mineral; *B*, cloth; *C*, cord; *D*, strand of fibers used in weaving cloth.

use on a large scale, you will see why only a few of the many fibers with which nature provides us are suitable for use.

Animal and vegetable fibers. The common fibers may be divided into animal and vegetable fibers. The animal fibers, wool and silk, are nitrogenous compounds. They may be distinguished from other fibers by the peculiar odor they give when burning. Heat readily affects them, even a moderate heat weakening them materially. They are attacked by alkalies, but withstand the action of dilute acids better than the vegetable fibers.

The vegetable fibers, cotton and linen, are made up of plant cells, and are largely cellulose. They can withstand quite a high temperature and the action of alkalies, but acids attack them and impair their strength.

Asbestos and glass fiber. Outside of these two main classes, we have such fibers as asbestos. This is a mineral that occurs in long white or greenish white fibers. These are very fine, but possess little cohesiveness, and are brittle. It is therefore difficult to make a fine asbestos thread. The fibers can, however, be woven into cloth, and will felt together into a paperlike sheet (Fig. 111).

Asbestos is a poor conductor of heat, is incombustible, and melts at a high temperature. The cloth woven from it is therefore suitable for the fire curtain of a theater to shut off the auditorium from the stage in case of fire, for firemen's gloves, for a protection to woodwork behind stoves, and for similar uses. The mineral is cheap, and asbestos paper, cloth, and sheets are used extensively.

Glass can be drawn out into fine thread and this has been used to weave cloth. Its commercial use is of course impossible, owing to its weight and brittleness.

Even spiders' web has been used as cloth material, but its cost and scarcity make the product only a curiosity.

Cotton. Cotton, an impure form of cellulose, is the most important fiber known. Our earliest records speak of its use. Herodotus tells us of its use in India, and the army of Xerxes was clothed in it. It is the product of the cotton plant that is grown so largely in our Southern States. The world production is enormous, the annual growth being about 18,000,000 bales.

(*a*) Cotton. (*b*) Wool. (*c*) Silk.

FIG. 112. — Textile fibers.

Under the microscope, it is seen to be a thin twisted ribbon, something like a dumb-bell in cross section (Fig. 112). The best grade is known as Sea Island cotton, valued for its long, strong fiber, which is used in the best grades of fine thread and for mercerizing. The average length of a cotton fiber is 1.6 inches and its diameter 0.00064 inch. In strength it comes between wool and silk, but it is not as elastic as either.

If the crude fiber is treated with dilute sodium hydroxide, dilute acid, water, and then with alcohol, all of the impurities are removed, and a soft, pure cellulose remains. This is the absorbent cotton of the household.

Cotton goods, if tightly stretched and treated with concentrated cold sodium hydroxide for a short time, and then well washed with water, gain about 30% in strength and acquire a luster like that of silk. The process is called mercerization, and the product, mercerized cotton. If the process is well done, mercerized cotton can scarcely be distinguished from silk.

Linen. The flax plant gives us the fiber called linen. Its use in cloth dates back even earlier than the use of cotton, for mummy cloths of linen have been found in Egypt that date from about 2000 B.C. The plant grows to a height of about 40 inches, and it is the stalk that yields the fiber. In addition we obtain linseed oil from the seeds.

The flax stalks are placed in stagnant water, where an active fermentation soon starts. This allows the linen fiber to be separated from the other plant products by mechanical means. The length of the linen fiber varies from three feet to a few inches. It is smooth, and has a somewhat silklike luster. It is much stronger than cotton. As it is a much better conductor of heat than cotton, it is cool when used as a dress material.

Wool. Wool is the hair of sheep. There are many varieties, depending on the variety of sheep from which it is obtained. Cashmere comes from the Thibet goat, mohair from the Angora goat, and alpaca from the llama. Not all of the dress goods having these names come from the proper source, as the peculiar characteristics can be imitated by common wool.

All wool fibers have an outer layer of flattened cells, the edges of which project outward, giving a saw tooth appearance under the microscope (Fig. 112). When the wool fibers are beaten together, this peculiar structure causes them to interlock, and felt. These scales are small and numerous. In mohair there are 2000 scales to the inch.

Wool fibers are from 1 to 8 inches in length. Wool is bleached with hydrogen peroxide, as chlorine attacks the fiber. Wool is readily soluble in sodium hydroxide, which gives us an easy way of distinguishing it from other fibers. It dyes easily and the colors are usually fast.

As wool is comparatively expensive, old worn woolen cloth is picked to pieces, woven again into yarn, and used with a proportion of new wool to make cloth. This "shoddy" has a short fiber, is weak, and will not wear as well as new wool. It is of course cheaper.

Much of the "all wool cloth" of commerce contains cotton. If it were sold for what it is, the cotton would not be objectionable, for the partly cotton cloth can be produced more cheaply than the all wool. It is not as warm as wool, and the two materials shrink differently, so that it must be used with care. It is a simple matter to detect the fraud, for if such a cloth is placed in hot dilute sodium hydroxide the wool will dissolve, while the cotton will remain (Fig 113).

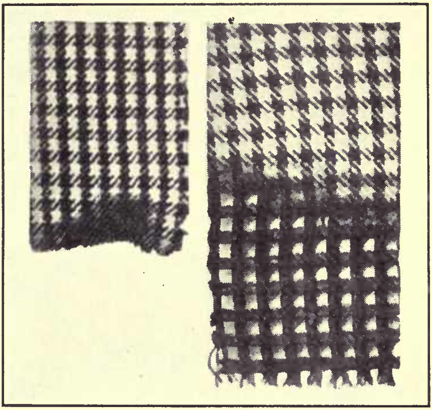


Fig. 113. — Testing for "all wool." The ends of two pieces of cloth were placed in hot dilute sodium hydroxide. The "all wool" is shown on the left; the mixed wool and cotton on the right.

Silk. Silk is obtained from the cocoon of the silkworm. The cocoons are heated to 70° C., which kills the worm. The silk is then unreeled, the silk fibers from several cocoons twisted together, and thus a silk thread is obtained. Only about 10% of the weight of the cocoons is obtained as thread. Five per cent is waste silk and broken threads used in spun silk, while the rest is waste.

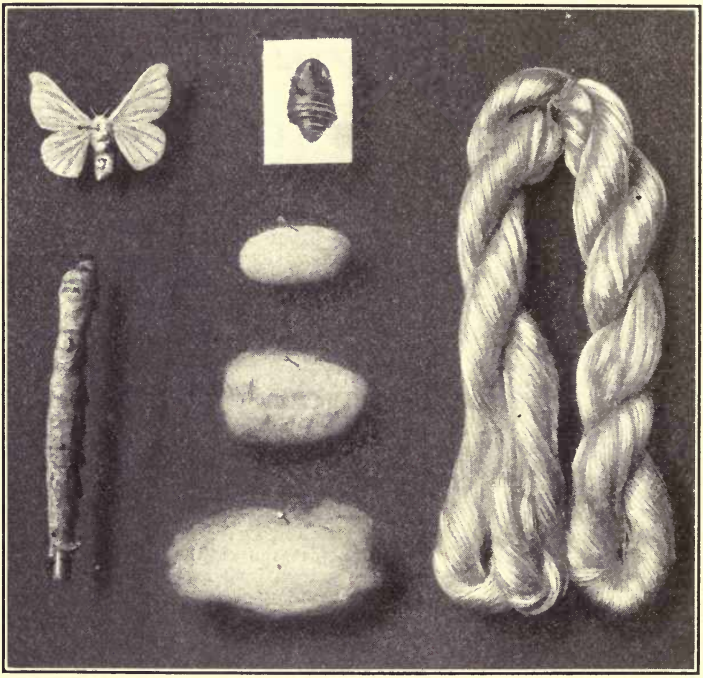


FIG. 114. — Silkworm, cocoons, chrysalis, moth, and skein of silk.

Silk thread has a beautiful luster, and a high tensile strength, 64,000 pounds per square inch. The average diameter of the thread is about 0.0005 inch. It is readily dyed. The fiber is round and smooth (Fig 112). It is easily soluble in concentrated hydrochloric acid, or in basic zinc chloride, which gives us an easy means of distinguishing the fiber.

Artificial silk. Owing to the beauty of silk, and its high cost, many attempts have been made to prepare an artificial silk. There are several products on the market that go by the name of artificial silk. They resemble it in luster, but have an entirely different composition. There are three main methods of preparing them.

Collodion silk is made by dissolving nitrated cellulose in a mixture of alcohol and ether, and squirting the solution through a small opening. The fine thread that forms is passed through water, which coagulates it.

Cuprate silks are made by dissolving cellulose in ammoniacal copper oxide. The solution is forced through fine glass tubes into water, which coagulates it. The copper is removed by a treatment with dilute acid and the thread dried.

Viscose silk is a solution of mercerized cellulose in sodium hydroxide and carbon disulphide. It forms a thick yellowish gelatinous mass, soluble in water. This solution is forced through fine openings into a solution of ammonium sulphate which coagulates the fine thread.

All of these artificial silks have the luster of silk, but are deficient in strength, especially when wet. They are difficult to dye, as when wet they are so weak that they are easily injured. They are used extensively as silk substitutes, as in the manufacture of braid, neckties, and artificial horsehair.

SUMMARY

Asbestos is a silicate that occurs in long fibers. It is used in making heat insulation cloth.

Cotton is almost pure cellulose. It is the most used fiber for cloth and thread.

Linen comes from the flax plant.

Wool is obtained from sheep. It is soluble in sodium hydroxide.

Silk comes from the silkworm. It is soluble in hydrochloric acid.

Artificial silk is made by dissolving cellulose, forcing the solution through small holes to produce a thread, and then coagulating the solution. It has the luster of silk, but is weak, especially when wet.

Exercises

1. How is felt made, and what causes the fibers to stick together?

2. What simple test will distinguish wool from cotton?

3. What is shoddy?

4. Would the fibers in rhubarb be suitable for weaving a cloth? Explain.

5. Paper can be made waterproof. Why is this material not a suitable dress goods?

CHAPTER XXVII

LAUNDRY CHEMISTRY

General methods of removing stains. To remove a stain from cloth, without injury to it, is often a difficult matter. The knowledge gained in your study of chemistry will help you, for when you know the properties of the substance causing the stain, you can often tell what means to use to remove it with certainty and safety.

The general methods of removing a stain are: to dissolve it, to absorb it, to bleach it, or to neutralize it. Many delicate colors are altered or destroyed by the chemicals used to remove stains. If a small piece of the goods can be had, it is well to try the effect of the reagent you expect to use on it, so that you may guard against using anything that will change the color of the fabric.

Removing stains by solution. The safest method to try first is solution. The solvent to be used will of course depend on the stain to be removed. Water will dissolve such substances as sugar, gum, and many salts. Always use cold water first, and plenty of it. A good way is to hold the stain under the cold water faucet, and allow the water to flow through it. After most of the stain has been washed out, warm water may be used to complete its removal. To remove a stain takes time. In bleaching cotton sheeting the manufacturer takes more than a day to complete the bleach. Do not expect water, or any other solvent, to remove a stain instantly. The addition of soap to the water will often help. The soap forms an emulsion with the staining material, and aids its removal.

Alcohol is another common solvent that will remove many stains caused by organic substances. Grass stains readily dissolve in alcohol.

Grease is soluble in benzine, naphtha, gasoline, and benzol. While all of these will readily dissolve the grease, their use is dangerous, as they are all inflammable materials. Fires have been caused by rubbing a silk waist in gasoline. The friction of the silk sometimes generates tiny electric sparks, which set fire to the gasoline, and a disastrous fire results. Carbon tetrachloride, CCl_4, will dissolve grease; it is cheap and noninflammable. It should be used instead of gasoline. Carbona is largely carbon tetrachloride.

Fresh paint is easily removed by carbon tetrachloride or turpentine. After paint has dried, it must be softened, by soaking it in amyl acetate or pine tar oil, before it can be removed. Varnish is a solution of gums in either alcohol or oil. If a spirit varnish is to be removed, alcohol will dissolve it. If an oil varnish is to be removed, gasoline will dissolve it. Ether and chloroform are other solvents that are sometimes useful.

When using any solvent, you must remember that you are dissolving the stain and forming a solution. If this solution is not removed from the fabric, you have not removed the stain, but spread it. If, starting at the center of the spot, you rub a grease spot with a little gasoline, the grease dissolves, the solution spreads, and you have a ring of grease. To avoid this, start at the edge of the spot, and work toward the center. Use plenty of solvent, and as this dissolves the grease, remove it, and replace it with fresh.

Removing stains by absorption. Grease spots can also be removed by absorption. Mix gasoline with talc or starch to a thick paste, and place this over and under the spot. Capillary action will draw the grease solution up into the talc. When dry, brush off the talc, and repeat if necessary. Blotting paper may be used in the same way. Place a piece of blotting paper over and under the spot, and press with a warm iron. The grease will be drawn into the paper.

Necessity of using a suitable solvent. The solvent must be suited to the material to be removed. You know that iron oxide is insoluble in all of the solvents that have been mentioned so far. It is then useless to try to remove iron rust with any of them. You must find a solvent for iron rust before you can remove it by solution. Acids will dissolve rust, but strong acids are likely to injure the cloth. An acid must be selected that will do as little harm as possible, and used weak.

The acid that we select depends, too, upon the fabric. Silk is soluble in hydrochloric acid. Obviously hydrochloric acid cannot be used on silk. A weak solution of oxalic acid may be used. On cotton or linen cloth, weak hydrochloric acid will remove the rust more quickly. After the acid has served its purpose, it must be carefully removed, or it will injure the cloth.

Removing acid stains. Since most anilin dyes are indicators, a drop of either acid or alkali will change their color. The remedy is to neutralize. If lemonade spilled on a silk skirt changes its color, a drop of dilute ammonia will restore the color, for the change produced by the lemon was due to the acid that it contained. In the same way, washing soda will alter colors. A few drops of a dilute acid will restore them. Do not forget to wash out the materials used.

Bleaching with Javelle water. Mildew cannot be dissolved. We must resort to another method to remove it. One simple way is to bleach it. A weak solution of Javelle water[1] may be used. The mildewed cloth is soaked in weak Javelle water, rubbed well, and allowed to remain until the mildew is oxidized. The cloth is then well washed. Many stains caused by anilin dyes can also be oxidized by the use of Javelle water. Chlorine must not be used on either wool or silk.

Bleaching with sulphur dioxide. Straw hats that have turned yellow may be bleached with sulphur dioxide. The hat is first scrubbed with soap and water to remove dirt, and, while still wet, hung up in an inverted barrel. A little sulphur is burned under the barrel. The sulphur dioxide produced will bleach the straw.

Bluing clothes. After clothes have been washed they commonly have a slight yellow tinge. They may be bleached by the action of dew and sunlight. If the clothes are spread on the grass, and allowed to remain some time, they will be bleached to a pure white. Most city dwellers find it impossible to do this, and bluing is resorted to. The slight blue tinge given to the clothes neutralizes their yellow tinge, and they seem to be pure white.

A number of different substances are used for bluing. Formerly indigo was used. As it is insoluble in water, a lump of indigo was tied in a cloth and swished through the water. This left indigo in suspension. If care was used, this was satisfactory, but, if the bluing water was allowed to stand, the indigo settled, and the clothes were streaked.

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[1] See the laboratory manual for its preparation.

Ultramarine is now used as a substitute for indigo. Its action is the same as that of indigo. Reckitt's blue belongs to this class of insoluble blues.

Prussian blue is used as a cheap substitute for ultramarine. This has the disadvantage that if the goods are not well rinsed before the blue is used, it will decompose and produce iron rust on the clothes. It is not as satisfactory as either indigo or ultramarine.

Soluble washing blues are largely aniline colors. They are easy to use, as they do not settle, and give the clothes a satisfactory color. Dyes are used that do not permanently stain the fibers.

The use of soap. Soap is used for ordinary cleansing of clothes. A good grade of laundry soap should not contain too large a percentage of water, and no free fat. The yellow laundry soaps are made in part from rosin. This is hardly an adulteration, as rosin soap is a detergent and lathers freely. It tends, however, to make the clothes yellow. The presence of free alkali is objectionable, as it is harmful to clothing and to the skin.

Just how soap aids in the removal of dirt is a disputed question. Its action is probably both chemical and mechanical. Its action is largely due to the fact that soap solutions will break up and emulsify the fatty substances that hold the dirt on the clothes.

Washing powders. Washing soda acts upon grease and softens hard water. A mixture of washing soda and powdered soap makes up the majority of the washing powders. It is usually cheaper to use washing soda and soap than these prepared powders. When crystals of washing soda are used, they should be dissolved in water and a little of the solution mixed well with the wash water.

How laundries wash clothes. An outline of the method used in a large laundry in cleaning collars and cuffs will show the differences between commercial and home methods. The collars are marked to identify them, and then placed in cold water for five minutes. This loosens the dirt, and dissolves some albuminous matters that would be coagulated if the collars were put into hot water first. This water is then run off, and hot water and soap added, also washing soda if that is necessary. The water is gradually heated to the boiling point. Twenty-five minutes are allowed for this " first suds."

The dirty water is run off, replaced by fresh, and a small quantity of soap added. A bleach made from chloride of lime is also added. This " second suds " is gradually heated to the boiling point.

After twenty-five minutes this second soapy water is run off, and the collars rinsed twice with hot water. Ten minutes is allowed for each rinse. A third rinse containing acid is then given. The object of the acid is to neutralize any alkali remaining in the clothes, as this would destroy the bluing used.

The collars are then blued, rinsed again, and starched. The starch used is a thin boiling, cornstarch, and is rubbed into the collars by machinery. The excess of starch is then rubbed off, first with the hand, and then with cheesecloth, and the collars dried. They are then dampened and ironed.

The main objection to such laundries is the bleach employed. This is often used too strong, so as to hasten the whitening of the clothes. As a result, the cloth is made very tender, and is easily torn.

SUMMARY

Stains are removed by solution, absorption, bleaching, and neutralization.

Laundry blues are indigo, ultramarine, and Prussian blue. Soluble blues are aniline colors.

Washing powders are mixtures of soap and washing soda.

Exercises

1. How would you remove a grass stain from a white skirt?
2. Why does the juice of grapefruit stain colored goods, and how would you remove the stain?
3. Why do you put clothes to soak in cold water?
4. What is " dry cleaning "?
5. How would you clean your straw hat?
6. Why are blued clothes sometimes streaky?
7. Are yellow soaps economical?

CHAPTER XXVIII

THE CHEMISTRY OF COOKING

Advantages of cooked food. It is possible to sustain life on raw food, but all mankind, including even the savage races, has found that cooking renders many foods more palatable.

Food is cooked for many reasons. The flavor may be improved, as in the case of meats. The tough fibers are softened, allowing easy digestion, as in the case of vegetables. Any parasites present are killed, thus making the food more healthful.

Cooking coagulates the protein present. Dry heat converts starch into dextrin and glucose. Moist heat swells the starch granules and causes them to break open. Sugars are changed to caramel, and fats decompose. All these changes cause alterations in the flavor of the food, and often produce savory odors that arouse the appetite.

How to cook protein. Boiling an egg is perhaps the simplest operation in cooking, and yet if you understand how to do this, you will understand the fundamental principles underlying the cooking of all meats. Albumin when heated to 158° F. coagulates. If the raw egg is put into boiling water, the outside layer of albumin quickly becomes hard and tough, while the inside of the egg is still raw. If, however, the egg is placed in cold water, and gradually heated, it becomes hot all through, the albumin coagulates throughout to a jellylike mass, and the egg is more palatable and digestible.

In boiling meats, if the meat is put into cold water, the albuminoids and extractives dissolve, and pass into the water. This makes the water rich, but leaves the meat dry and tasteless. We should therefore put the meat into boiling water. This will coagulate the albumin on the outside and prevent the escape of the juice. If we wish to extract the juice, as in making soups and beef tea, we should cut the meat into small pieces, put these into cold water, and heat them slowly.

Simmering. Simmering, which is cooking at a temperature of 185° F., makes meat tender. The meat is first put into boiling water to coagulate the outside, and then the temperature lowered. This keeps the fibers tender, but the meat must of course be cooked for a longer time.

Fried steak has ruined many a farmer's stomach, but the fault is not in frying, but in the way in which it is done. If a steak is put into a cold frying pan, with some fat, and then heated, the fat is drawn up between the fibers of the meat. Each individual fiber is covered with a layer of fat, and a greasy, dry, indigestible mass results. If the steak had been put on a very hot frying pan, the outside ends of the fibers would have been seared, and a result almost equal to broiling obtained.

Deep-fat frying. Frying by immersion in deep, hot fat is a common method of cooking many foods. Fat does not boil, but when heated to a temperature considerably above the boiling point of water, it smokes. It is then almost at its kindling temperature, and care must be taken that it does not catch on fire. As smoking fat is far above 212° F., water thrown into it instantly changes into steam. This causes a slight explosion, and hot fat is thrown around. It is for this reason that whatever we wish to fry should be dry.

Fried foods are greasy because cooks do not keep the fat hot enough. When, as in the home, a small quantity of fat is used, to throw a large handful of potatoes into it cools the fat so much, that, instead of a crust forming on the outside of the potatoes, which prevents the entrance of the fat, the grease sinks in between the starch grains. Naturally a greasy, indigestible mass is the result.

Rule for cooking meat. There are only two things to remember in cooking meats. If you wish to keep the juice in, heat quickly to coagulate the outside. If you wish to extract the juice, heat slowly, and keep the temperature below 185° F.

Cooking vegetables. The chief advantage gained in cooking vegetables is in making them more digestible. Heating swells the starch granules and they burst, tearing apart the cellular structure in which they are imbedded. The digestive juices are then able to get at the starch and digest it.

Vegetables are best placed in plenty of boiling water if we wish to retain their flavor. When we wish to extract part of the flavor, as in the case of onions, they are best placed in cold water, and then brought slowly to a boil.

A fallacy of cooks. One very common mistake of the cook is in thinking that hard boiling heats water to a higher temperature than slow boiling. You know that water boils at 212° F. Whether it is boiling furiously, or is only just boiling, makes no difference. It is impossible to raise its temperature. To boil potatoes hard, then, is to waste fuel.

SUMMARY

Cooking food makes it easier to digest, more palatable, and safe.

Cooking meats: To keep meat juices in, place the meat first in boiling water. To extract meat juices, place the meat in cold water and heat slowly.

Simmering is cooking by heating in water at a temperature of 185° F.

Temperature of boiling water. Water has the same temperature, whether boiling slowly or furiously. In an open vessel at sea level it is 212° F.

Exercises

1. How would you hard-boil an egg?

2. How would you prepare beef tea?

3. Why do thin pieces of potato swell when dropped into very hot fat?

4. Why is the meat left in preparing soup so tasteless?

5. In cooking vegetables should they be put first into hot or cold water? Why?

6. What causes greasy doughnuts?

7. Why is toast rather than bread given to invalids?

CHAPTER XXIX

CALCIUM AND ITS COMPOUNDS

The metal calcium. The element calcium is a hard, silvery metal, that has no commercial use. Calcium is never found free in nature, for, like sodium, it oxidizes on exposure to moist air. It is easily prepared by the electrolysis of the fused chloride. Its compounds are important, forming about one fifteenth of the bulk of the earth. You are already familiar with some of them, as lime and marble.

FIG. 115. — Limestone, showing its shell origin.

Calcium carbonate. Calcium carbonate, $CaCO_3$, occurs in the most varied forms, sometimes beautifully crystallized, and sometimes as an amorphous rock. Marble, limestone, and the shells of shellfish are almost entirely calcium carbonate. Chalk is the shell remains of a microorganism and is almost pure calcium carbonate. When washed and purified, it is called *whiting*, and is used for calcimine. Whiting mixed with linseed oil forms putty. Pearls are calcium carbonate, as is also coral. When pure, it occurs in large, transparent crystals, called Iceland spar and dogtooth spar.

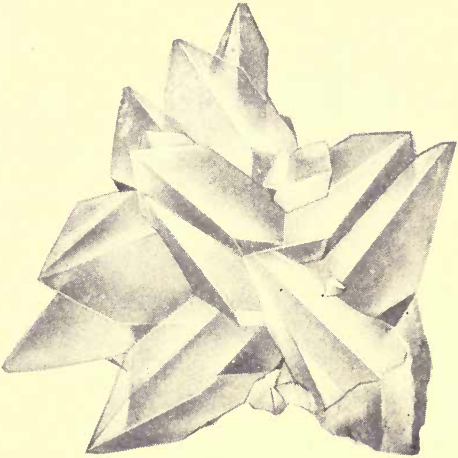
Fig. 116. — Dogtooth spar.

All varieties of calcium carbonate are easily acted on by acids, carbon dioxide being given off.

$$CaCO_3 + 2\ HCl \rightarrow CaCl_2 + H_2O + CO_2 \uparrow$$

This serves to distinguish it from other white rocks.

Marble. Marble is insoluble in water, and it is an interesting question how such animals as corals can obtain the calcium carbonate necessary to form their skeletons. If you have ever noticed the marble slab under the faucet of the soda fountain, you will have seen that it is very badly worn, even though the rest of the soda fountain is new. These slabs have to be renewed every year. This gives us a clue. Carbon dioxide when it dissolves in water forms an acid called carbonic acid, and marble is soluble in this acid, forming calcium bicarbonate.

$$H_2O + CO_2 \rightarrow H_2CO_3$$
$$H_2CO_3 + CaCO_3 \rightarrow CaH_2(CO_3)_2$$

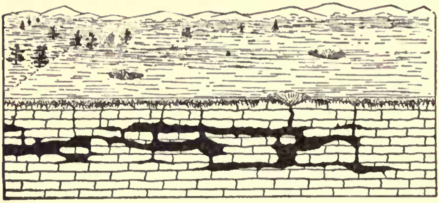

FIG. 117. — How caves are formed by water containing carbon dioxide.

When it rains, the rain water dissolves carbon dioxide from the air, and forms a weak solution of carbonic acid. This sinks into the ground, and, if it runs over marble, will dissolve it, forming a solution of calcium bicarbonate. This runs to the sea, and it is from this source that the coral animals and the shellfish obtain the calcium carbonate necessary for their growth.

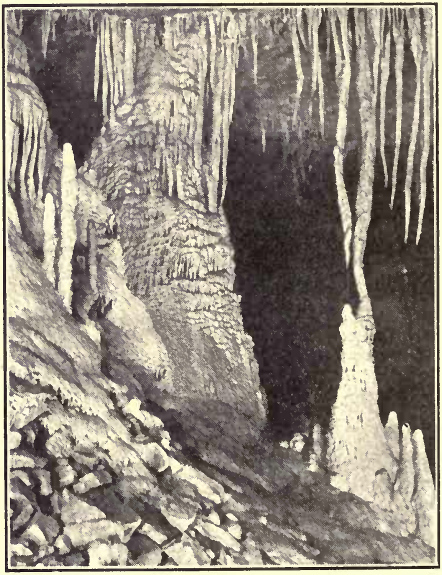

FIG. 118. — Stalactites and stalagmites in Mammoth cave.

How caves are formed. This dissolving of marble or limestone leaves cavities in the ground. These are caves, such as the Mammoth cave of Kentucky. As marble varies in hardness, some parts of it will be more easily dissolved than others, hence the winding passages, the large rooms, and low channels found in caves.

If water containing calcium bicarbonate trickles over the roof of a cave, and evaporates, calcium carbonate is left. As water continues to flow in and evaporate, a stone "icicle," called a *stalactite*, forms. A slight wind blowing through the cave will force the water to one side of the stalactite, or a grain of sand will divert its course. In such ways the most curious stalactitic forms arise.

If more water trickles down the stalactite than can evaporate, some drops to the floor of the cave. Here it forms a pool that slowly evaporates and deposits a mound that slowly grows upward. This is a *stalagmite*. When a stalactite and a stalagmite meet, they form a column.

As the water occasionally contains iron and mud as impurities, these growths are sometimes white, and sometimes brown, pink, or dark gray. It requires only a little imagination to see in these grotesque growths the Organ, the Bishop's Chair, and other fanciful objects that the guide points out to you.

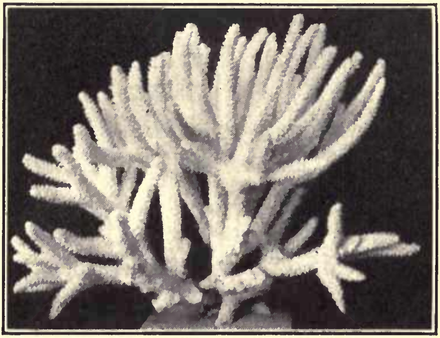

Fig. 119. — Coral.

Coral rock. Coral animals cannot live out of water, so when they have carried their skeletons to within a few feet of the surface of the ocean they stop building. Storms break off the top portion of the coral rock, grind it to powder, and this fills up the cracks of the coral. Floating vegetable matter drifts on to the coral and is caught, shells and seaweeds are washed up, and finally, as a result of all these agencies, a land area is formed above the surface of the ocean. Birds then use this as a resting place and deposit seeds there, or, perhaps a floating coconut strands on the shore. Soon vegetation starts growing, and we have a verdant coral isle.

If you have been in Florida and noticed the rock (coquina) of which St. Augustine is largely built, you know that it is composed of broken shells. In time, these broken shells become compact, and limestone results. This completes the cycle, that will recommence when rain water falls on the limestone and again dissolves it. All marble and limestone deposits are the remains of life. In many specimens of limestone you can make out faint traces of the shells from which it was built.

Gypsum. Gypsum is one of the common white minerals found in many parts of the United States. It is so soft that it can be scratched by the finger nail. When pure it is white, but it is often colored red, yellow, or brown, by the presence of small amounts of impurities. Chemically it is calcium sulphate, $CaSO_4$, containing two molecules of water of crystallization, $CaSO_4 \cdot 2\,H_2O$. When heated, gypsum loses most of this water of crystallization and is converted into amorphous calcium sulphate. This, when mixed with water, combines with it, crystallizes, and forms gypsum once more.

Plaster of Paris. Anhydrous gypsum is ground to a fine powder called plaster of Paris, because it was first made near the city of Paris. This is used to make plaster objects by mixing it with enough water to form a thick mud, and pouring into molds. The water combines with the calcium sulphate, forming gypsum, the mass sets or becomes hard, and a plaster cast is thus obtained. Large amounts of plaster of Paris are used in making plaster casts, stucco, the final finish of walls, and the outside of such temporary buildings as are often erected at seaside resorts.

Gypsum is slightly soluble in water, one part of gypsum requiring 500 parts of water to dissolve it at ordinary temperatures. Since gypsum is a common mineral, it follows that in many parts of the country the spring waters will contain gypsum in solution.

Making of quicklime. When heated to a red heat, marble decomposes, and gives off carbon dioxide, leaving calcium oxide, called quicklime.

$$CaCO_3 + \text{heat} \rightarrow CaO + CO_2 \uparrow$$

The making of quicklime is carried out commercially by heating marble or limestone in large furnaces called lime kilns. The product is a white solid used in making mortar. The kiln is filled with limestone. The hot gases from the furnace constantly mingle with it. At about 1800° F. carbon dioxide escapes from the limestone and lime results. This is drawn from the bottom of the kiln from time to time.

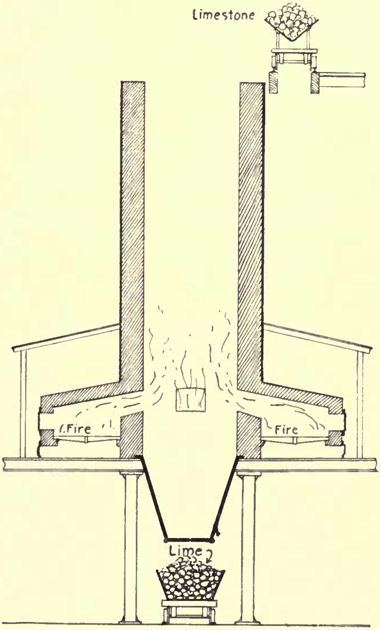

FIG. 120. — A lime kiln.

Slaked lime. When water is added to quicklime, it swells up, and is changed to slaked lime or calcium hydroxide.

$$CaO + H_2O \rightarrow Ca(OH)_2$$

A large amount of heat is developed during this chemical change. It is not uncommon for a barge carrying quicklime to catch on fire, if, by chance, a little water reaches the lime. It is a dangerous chemical and must be stored with care.

Self-heating canned soups have been made by taking advantage of this heating effect. A sealed pint can of soup is placed in a quart tin can, and the space between the two is filled with quicklime. The outer can is then sealed. When the soup is to be used, holes are punched in the outer can and water poured in. The lime slakes, and the soup becomes hot. This method of heating is of use to balloonists, who dare not use a fire, lest the gas contained in the balloon catch on fire.

Mortar and plaster. Mortar is made by mixing quicklime, water, and sharp sand. This mixture is placed between bricks to hold them together. The calcium hydroxide that is produced absorbs carbon dioxide from the air, and changes back to marble.

$$Ca(OH)_2 + CO_2 \rightarrow CaCO_3 + H_2O$$

In time the mortar becomes as hard as the bricks themselves.

The plaster used to cover the laths nailed on walls is mortar to which some hair has been added to increase its tenacity. Newly built houses are unhealthful because both the mortar and plaster, although seemingly dry, are giving off large quantities of water. This makes the house damp. Then, too, the hair in the damp plaster slowly decomposes, and this is unpleasant.

Limewater. Calcium hydroxide is slightly soluble in water. The filtered solution is called *limewater.* It is used as a test for carbon dioxide, and is mixed with milk to make it more digestible for invalids. It must be protected from the air, as otherwise it absorbs carbon dioxide and becomes worthless.

As calcium hydroxide is a base, and is cheap, chemists use it in large amounts to neutralize acids. It is also used in the manufacture of sodium hydroxide, in bleaching powder, and in purifying illuminating gas.

Calcium phosphate. Calcium phosphate, $Ca_3(PO_4)_2$, is the main mineral constituent of our bones. It occurs as a rock in Canada, Florida, South Carolina, and other places. The phosphorus in this rock is valuable as a fertilizer, but must be changed into a soluble form before plants can use it. This is done by treating the phosphate with sulphuric acid, when calcium superphosphate is formed.

$$Ca_3(PO_4)_2 + 2\ H_2SO_4 \rightarrow 2\ CaSO_4 + CaH_4(PO_4)_2$$

This superphosphate is mixed with some nitrogenous material and forms a commercial fertilizer.

Bleaching powder. Chloride of lime, bleaching powder, $CaOCl_2$, is made by passing chlorine gas over layers of powdered slaked lime. It is a white powder that smells feebly of chlorine. The addition of an acid sets free chlorine, as you have seen in your laboratory work.

$$CaCl(OCl) + H_2SO_4 \rightarrow CaSO_4 + HCl + HClO$$
$$HCl + HClO \rightarrow H_2O + Cl_2$$

It is used as a source of chlorine and for disinfecting.

Hard waters. When water containing a soluble calcium salt is brought into contact with soap, a white precipitate of lime soap is formed. Such water is called hard water. There are two kinds, temporary and permanent hard water.

Temporary hard water contains calcium bicarbonate. It may be softened by heating, which drives out the carbon dioxide, leaving calcium carbonate as a white precipitate.

$$CaH_2(CO_3)_2 + \text{heat} \rightarrow CaCO_3 \downarrow + CO_2 \uparrow + H_2O$$

When large quantities of temporary hard water must be softened, heating is too expensive. Instead, lime is added, and the water allowed to stand until clear. Exactly the right amount of lime must be added, as otherwise, the excess of lime will itself make the water hard. This method of softening can only be carried out successfully by a chemist.

$$CaH_2(CO_3)_2 + Ca(OH)_2 \rightarrow 2\ CaCO_3 \downarrow + 2\ H_2O$$

Permanent hard water contains calcium sulphate, or sometimes magnesium sulphate or chloride. It can be softened by adding sodium carbonate.

$$CaSO_4 + Na_2CO_3 \rightarrow CaCO_3 \downarrow + Na_2SO_4$$

Disadvantages of hard water. Hard water used in the home has several disadvantages. It deposits a fur on the inside of teakettles, and this deposit, being a nonconductor of heat, renders it difficult to heat the water.

In washing, large amounts of soap must be used to precipitate lime soap, and this is waste, not only on account of the soap that is consumed, but because the lime soap is sticky and soils the clothes.

Hard water is also objectionable in industries. It forms a scale in boilers, and interferes with many chemical processes.

How hardness is measured. The hardness of water is always given in degrees. One degree is one grain of calcium carbonate, or its equivalent in other salts, in one gallon of water (58,333 grains). The higher the degree of hardness, the worse the water for domestic use. One hundred gallons of water, of ten degrees hardness, requires about one pound of soap to soften it.

SUMMARY

Calcium is a silvery white metal that oxidizes on exposure to moist air.

Calcium carbonate exists as marble, limestone, coral, shells of shellfish, pearls. Gypsum is calcium sulphate. Heated and ground, it forms plaster of Paris.

Quicklime is calcium oxide. Slaked lime is calcium hydroxide.

Mortar is a mixture of quicklime, sharp sand, and water.

Bleaching powder is made by passing chlorine over powdered calcium hydroxide.

Temporary hard water contains calcium bicarbonate. It is softened by boiling or by adding slaked lime.

Permanent hard water is water containing calcium sulphate, or magnesium sulphate or chloride. It is softened by adding sodium carbonate.

Hardness of water is measured in degrees. One degree is one grain of calcium carbonate, or its equivalent in other salts, in one gallon of water.

Exercises

1. How would you obtain a plaster replica of a medal?

2. How should putty be stored, and why?

3. Why does putty harden?

4. If you breathe through limewater, it first becomes cloudy and then clears. Why?

5. How could you distinguish between a temporary and a permanent hard water?

CHAPTER XXX

DYES AND DYEING

Usefulness. Many of us have in our home materials that have been discarded, not because of wear, but because the color has faded, or no longer goes with the furnishings of a room. In many such cases it is possible with a little care to dye the article at home so that it again becomes useful. Dresses, satin slippers, draperies, curtains, straw hats, and feathers, all may be done as well at home as by a professional, if only care and a little chemical knowledge are used.

Preliminary precautions. Before dyeing anything it must be clean. This means more than a mere surface cleanliness. If a grease spot has been imperfectly removed, it may not show on the goods, but it will prevent the even action of the dye and cause a spot. Perspiration stains, fruit stains, rust spots, all must be removed if we are to have a perfect result. It is well to remove any buttons, especially if of metal, and any bead or metal trimming, as these interfere with some of the dyes used. It is impossible to dye a light color over a dark, and in general the colors produced are more satisfactory on white cloth. It is well therefore to remove as much of the original color as possible before attempting to re-dye. Often boiling in two changes of water will discharge much of the original color.

Care necessary. First select a suitable dye. This means not only to pick out a dye of the color you wish, but also one that will be suitable for the material used. When you have studied the later paragraphs of the chapter, you will understand how to pick the dye, depending on whether the material is vegetable or animal fiber. Special care must be used in mixed goods. Many dyes do not take on cotton as well as they do on wool, so that these dyes used on a mixture of wool and cotton would produce a mottled effect.

Having selected the dye, dissolve it in a little water (assuming it to be one of the common aniline dyes), and filter through cloth to remove any undissolved particles. Dilute the solution to the required volume in a large vessel that will stand heat. Do not use iron pots, as any rust will prevent success, and the iron will discolor many dyes. Heat the dye until almost boiling and then immerse the goods, having first wet them thoroughly. This is important and is often overlooked. If the goods are not kept constantly in motion, different parts of the cloth will receive different amounts of the dye, and the color will be uneven, or not " level." Therefore with the aid of two sticks lift the goods and move them about during the time required to complete the dyeing. Most dyes take best at the boiling temperature, but this does not mean that it is necessary to boil them furiously; simply keep them at the boiling temperature. The shade of the goods after drying is generally lighter than while wet, and some experience is necessary to match shades exactly. For this reason, and because of the uncertain character of the fiber in some goods, it is well to dye a sample first, noting the time, etc., and then treating the main bulk of goods in accordance with the experience gained.

To finish, rinse in cold water until the wash water is colorless, then dry and press. Remember especially to enter the goods wet, not to hurry the operation, not to try to dye light colors over dark, and most important of all to use a suitable dye.

Obtaining dyes. The coloring power of the artificial dyes is so great that only small quantities of them are needed. They can be bought by the ounce from chemical dealers costing from 25 to 75 cents an ounce. The small packages of the dyes on sale under various trade names are satisfactory, but they are more costly in proportion to the amount of dye received.

Fast colors. One annoying thing about colored goods is their tendency to change color. The fastness of colors is measured by their resistance to the different changing agents. Thus a color may be fast to light, but not fast to washing, or fast to washing but not fast to perspiration. The only way to be sure that any given piece of goods will meet your requirements is to try it.

An easy way to try fastness to light is to cut a sample in two, and put one half in bright sunlight, while the other half is kept in a book away from the light. After a few days compare the color. It is easier to get quick results with this test if you will take a small sample of the dye itself, dissolve it in water, and smear a little on paper. The color smear should be dark at one end, and very light at the other. Cut in two as before and expose to light. As one end of the smear is so light, it will show a change of shade much sooner than will a dyed cloth containing a considerable amount of the dye.

Your chemical experience will suggest to you similar ways in which you can test any dyed material for any desired quality of fastness. Many of the artificial dyes are of exceedingly good quality in every way, and it is possible to find a dye that will give almost any color on any material and that will be fast enough for any practical end.

Natural dyes. Formerly all dyeing was done with the aid of indigo, logwood, cochineal, etc. Most of these have been superseded by artificial dyes. Logwood, however, still holds its own in wool dyeing. Many of these colors are dyed by the aid of mordants, and our next study will be to see just how a dye that has no natural affinity for a fiber can be made to stick to it.

Lakes. The addition of ammonia to a solution of an aluminium salt yields a white gelatinous precipitate of aluminium hydroxide. This has an affinity for dyes, so that if a solution of litmus is shaken with the precipitate the color and the aluminium hydroxide form a loose chemical combination called a *lake.* The lakes used in painting, such as carmine lake and rose lake, are made in this way. These lakes may have iron, tin, chromium, or other metals substituted for the aluminium.

Aluminium as a mordant. Cotton has no affinity for logwood dye. If we soak cotton in a logwood solution, it will be stained, but on washing, the stain will almost entirely disappear.

The color is not fast. Let us, however, first soak the cotton in alum, and then in ammonia. This will precipitate aluminium hydroxide on the fiber. If the cotton is now soaked in logwood, the color will be absorbed, not by the cotton but by the aluminium hydroxide, forming a lake. The color will then be fast. Materials used in this way to fasten colors on fabrics are called *mordants,* and the dyes are known as mordant dyes. Salts of tin and iron and chromium are largely used, as well as tannic acid.

Logwood is one of the mordant dyes. It gives a good cheap black, and is used extensively on wool and silk. By using tin as the mordant, silk may be weighted so that the silk will weight 400 times as much as the raw material. Such weighting as this, however, is objectionable, as the silk will crack and wear poorly.

Indigo. Indigo blue, or indigotin, as the color is called, was formerly obtained from the indigo plant. It is now made synthetically in large amounts and the natural product is disappearing from the market. The color is insoluble in water, but when reduced becomes soluble, in which form it is colorless. Before dyeing, the indigo is reduced with acid sodium sulphite, $NaHSO_3$. This makes the indigo both soluble and colorless. The cloth is then entered and the dye made to soak into the fibers. The cloth is then removed and exposed to the air. The indigo reoxidizes and the indigo blue develops. It is an exceedingly fast color, and is extensively used for blue serges.

Artificial dyes. The artificial dyes come under many chemical classes, and we can consider only a few of them. The ones given below constitute some of the more important, and will illustrate the subject.

Acid dyes. The commercial dyestuff is usually an alkali or calcium salt of the color acid. When used with wool, the wool acts as a base, setting free and combining with the color acid, and forming an insoluble compound. Alkalies will render the color soluble and remove it, hence fibers dyed with these colors should not be washed with washing compounds containing alkalies. They are generally fast to light, and hence are useful in dyeing such things as feathers that do not require washing. Seventy-five per cent of the wool is dyed with these dyes in an acid bath. Acid dyes are also used with silk, but not for cotton. They are important in the dyeing of jute. The process of dyeing on wool is simple. The dye, together with 15 per cent Glauber's salt and 3 per cent of sulphuric acid, based on the weight of the material that is to be dyed, is placed in water heated to 140° F. when the goods are entered. The bath is then heated to boiling and the boiling continued for about three quarters of an hour. They are then removed and rinsed.

Basic colors. These colors are characterized by their bright or even gaudy colors. They are not very fast to light, but are to washing. They are basic in character, hence the name. They have an affinity for wool and silk, but not for cotton. They are used largely for silk. If we wish to use them with cotton, we must first mordant the fiber, using tannic acid.

Direct or substantive colors. The name direct is given to this class because they will dye directly on cotton. All vegetable fibers absorb them, as do the animal fibers, but for them the acid colors are preferable. The direct colors are used extensively in union goods, that is, mixtures of cotton and wool, and cotton and silk, as the dye will color both fibers. Glauber's salt is usually added to the bath to diminish the solubility of the dye and aid its deposition on the fiber. A concentrated bath is better than a dilute one. A treatment with a solution of copper sulphate after dyeing improves the fastness of the color. They vary much in fastness, being generally not fast to washing on cottons and faster on wools.

Bleeding. Many dyes, and the direct dyes on cotton in particular, are subject to a trouble called *bleeding.* If you will take a skein of cotton dyed with rhodamine or some other direct dye, twist it with a white skein, and then boil the two in water, you will find that at the end of an hour both skeins will be colored. The color has bled from one to the other. This aids level dyeing, but causes the color to run in washing.

Sulphur colors. These are known as sulphur colors, both because they contain sulphur and because sodium sulphide is used in the dye bath. They are fast to light, to acids, and to washing. They are only suitable for vegetable fibers, as the strong alkali used in the dye bath attacks the animal fibers. The colors are generally dull. A sample dye bath would be: dye 1–20 per cent of the weight of the material to be dyed; sodium sulphide 1–4 times the weight of the dyestuff; Glauber's salt 20–50 per cent; and soda ash 5–10 per cent of the weight of the material. Enter goods just below the boiling point, and boil for one hour. The colors are all insoluble in water, but are soluble in sodium sulphide.

Alizarin colors. The alizarin or artificial mordant colors are always used with mordants, usually chromium, aluminium, or iron. They are very fast to both light and washing. Alizarin is the most important dye of the group.

Miscellaneous colors. There are many other classes of colors, and many that do not fall into any special class. One of the most important of these is aniline black, produced by the oxidation of aniline hydrochlorate. There may be in your home a table that you would like to use for a home laboratory bench. To be suitable for this use it should be finished in such a way that chemicals will not readily act on it, and that it looks well, and is easily kept in condition. You may gain all these ends by staining it with aniline black.

First see that the wood is clean, and that no trace of a former finish is left. Then apply two coats of a solution composed of 5 ounces of copper sulphate, and 5 ounces of potassium chlorate, dissolved in 2½ pints of water. This solution should be applied hot, the second coat to follow as soon as the first is dry. When these are dry, apply a solution consisting of 6 ounces of aniline chloride crystals dissolved in 2½ pints of water. This solution should also be applied hot. When all is dry, rub in thoroughly raw linseed oil, using a cloth so as to secure a thin coating. Rub thoroughly and hard in order to bring out a good polish. Finally, wash with hot soapsuds. This is the formula used in many high schools to finish the laboratory table tops black.

SUMMARY

Material to be dyed must be clean, must be stirred while in the dye bath, and dyed hot.

Indigo, logwood, and cochineal are natural dyes.

A lake is the combination of a metallic hydroxide with a color.

A mordant is a substance used to make a color stick to a fiber for which it has no affinity.

Acid dyes are chiefly used for wool. Commercially they are the alkali or calcium salt of the color acid. Examples: acid magenta, patent blue, scarlet 2 R.

Basic dyes are used mainly for mordanted cotton. Examples: rhodamine, phosphine, magenta, bismarck brown.

Direct dyes are largely used in dyeing cotton and union goods. Examples: primulin, benzo orange R, Congo red 4 R.

Alizarin dyes. Very fast. Alizarin the most important.

Aniline black is made by the oxidation of aniline hydrochlorate.

Colors are tested for fastness by exposure to light, acid, or whatever special quality is desired.

Exercises

1. How would you dye an ostrich plume yellow? What class of dyes would be suitable?

2. A home-dyed dress showed after pressing large rings of a lighter color. What was one possible source of the trouble?

3. Could you make a red ink from a red dye? How?

4. What dyes would you use to dye raffia a bright red? Would the color be fast to light?

5. Why cannot a gray be dyed over a green?

6. There are many more shades of colored goods than there are dyes. How are they produced?

CHAPTER XXXI

SOME COMMON CHEMICALS

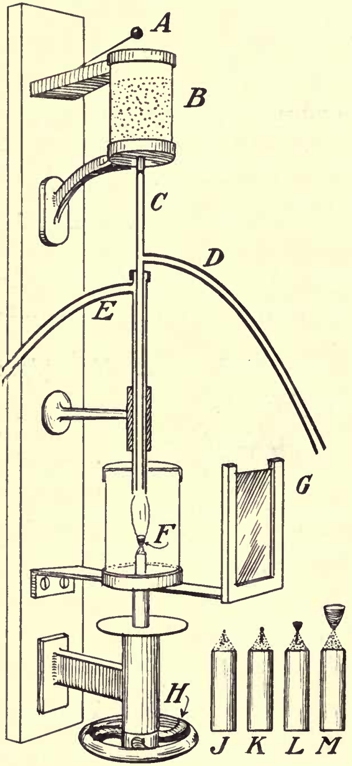

FIG. 121. — Making artificial rubies. The aluminium oxide is carried from *B* a little at a time through *C*. In passing through the oxyhydrogen flame it is fused and forms the *boule* *F*; *J–M* show supports with boules in various stages of completion.

Reference books needed. In an elementary chemistry it is impossible to consider all, or even a majority of, the common chemicals. All that such a chemistry as this can do is to start you on the road of chemical knowledge, and teach you how and where to look for additional information. When you wish to know about some chemical that is not studied in this chemistry, go to your library. The librarian will show you how to use the card index, and after a little practice you will see how easy and helpful the use of reference books is.

Some chemicals, however, that we have not studied are so important, that, although we cannot spare the time to study them in detail, a few facts about them should be known.

Aluminium oxide. Aluminium oxide, Al_2O_3, occurs in nature as the mineral *corundum.* This, when pure, forms transparent, hard crystals, that when cut are a good substitute for the diamond. When colored by the addition of small quantities of impurities, corundum forms the *ruby* and the *sapphire.* By melting aluminium oxide, mixed with small quantities of metallic oxides, it is possible to produce synthetic rubies and sapphires that are the equals of the natural in beauty (Fig. 121).

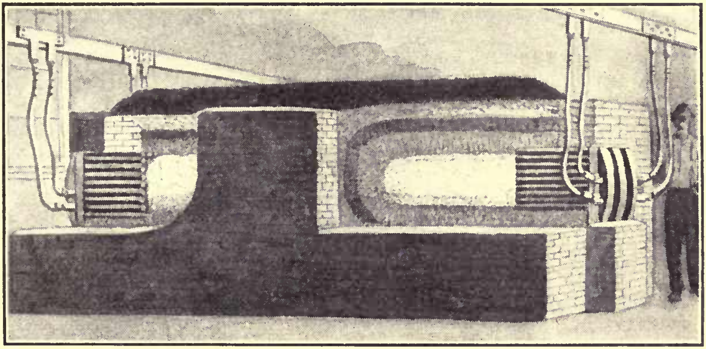

FIG. 122. — Carborundum furnace.

Much of the natural aluminium oxide occurs mixed with iron oxide in a black rock called *emery.* This is used as an abrasive. Aluminium oxide can be fused in the electric furnace, when it forms an exceedingly hard mass, that when broken up is called *alundum.* This is broken to a powder and used as an abrasive.

Abrasives. An abrasive is a material used in grinding and polishing hard substances. Sandstone and emery are two natural abrasives. Alundum is an artificial abrasive. By heating sand and coke in an electric furnace carbide of silicon is formed.

$$SiO_2 + 3\,C \rightarrow SiC + 2\,CO$$

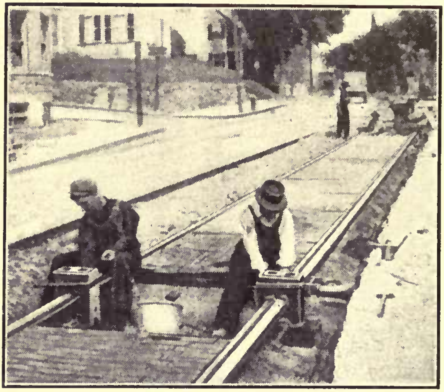

FIG. 123. — Thermit welding. Preparing molds about rail joints.

This silicon carbide is called *carborundum* and forms beautiful crystals, iridescent on the surface. It is so hard that it is widely used as an abrasive, many different kinds of polishing and grinding wheels, powders, and stones being made from it.

Thermit. A mixture of powdered aluminium and iron oxide, called thermit, burns with the production of an intense heat. At the same time the iron oxide is reduced to metallic iron and melts.

$$2\,Al + Fe_2O_3 \rightarrow 2\,Fe + Al_2O_3$$

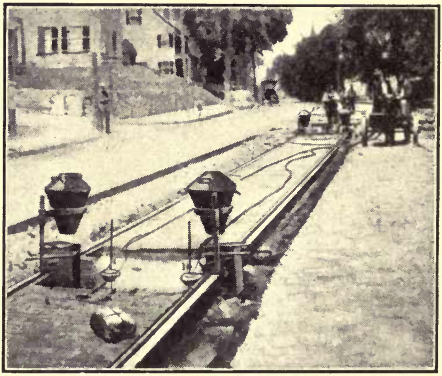

FIG. 124. — Thermit welding. Thermit in crucibles over the molds.

By igniting thermit in a crucible having a hole in the bottom that can be closed, placing the crucible over a broken iron article that is to be repaired, and then allowing the fused iron produced to flow out upon the break, the two pieces of iron can be welded together. In this way large broken articles can be easily and cheaply repaired.

By mixing aluminium with the oxides of other metals, as chromium, and igniting, the metallic oxide can be reduced and the pure metal obtained. This is a satisfactory method of obtaining such metals as manganese and chromium.

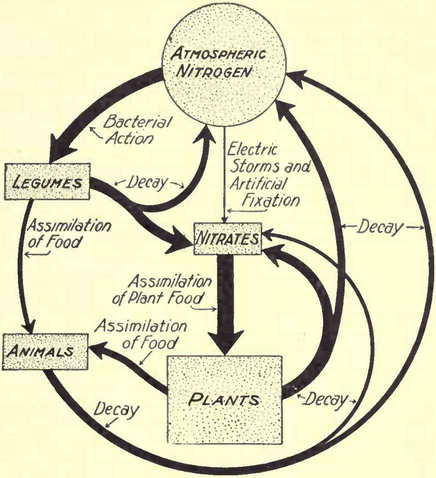

Fig. 125. — Nitrogen cycle. (From Blanchard and Wade's *Foundations of Chemistry*.)

Alums. The sulphate of aluminium forms loose chemical compounds with many other sulphates. These double sulphates are known as *alums*. Potassium alum, $K_2SO_4 \cdot Al_2(SO_4)_3 \cdot 24\,H_2O$, sodium alum, and chrome alum are common alums. They are used in dyeing, in water purification, and in baking powders.

Nitric acid and nitrates. Nitric acid is prepared by the action of sulphuric acid upon a nitrate. Sodium or potassium nitrate is commonly used. The mixture is heated in a retort, when the nitric acid distills over.

$$NaNO_3 + H_2SO_4 \rightarrow NaHSO_4 + HNO_3$$

Nitric acid is a colorless, heavy liquid having a sweetish smell, and is very corrosive. It is a strong oxidizing agent. Recently large quantities have been made by blowing air through electric arcs. The intense heat of the arc causes the nitrogen and oxygen of the air to combine, forming oxide of nitrogen. This, when dissolved in water, forms nitric acid. The process is important because this synthetic acid is used to prepare the nitrates that are indispensable in fertilizers.

This process is another illustration of the usefulness of chemistry to industry. Fertilizers must contain nitrogen in such a form that plants can assimilate it. The substance commonly used has been sodium nitrate. The supply of this Chili saltpeter (so called because it comes from Chili) is almost exhausted, and if it were not possible for the chemist to make synthetic nitric acid, the shortage of nitrates would soon cause very serious trouble to the farmer.

A mixture of nitric and sulphuric acids when run into glycerin changes it to a nitrate. This is *nitroglycerin,* the violent explosive.

$$C_3H_5(OH)_3 + 3\ HNO_3 + (H_2SO_4) \rightarrow C_3H_5(NO_3)_3 + 3\ H_2O + (H_2SO_4)$$

Absorbing this in ground wood, or some other porous material, gives *dynamite.* Some oxidizing material, as a nitrate, is also usually added. Most of the other high explosives owe their power to their being nitrates. *Guncotton* is cellulose hexanitrate $[C_{12}H_{14}O_4(NO_3)_6]_x$, prepared by treating cotton with a mixture of nitric and sulphuric acids. In making nitroglycerin and guncotton the sulphuric acid takes no part in the chemical change, but is used as a dehydrating agent. A lower cellulose nitrate is used in preparing collodion.

Phosphorus. When a mixture of calcium phosphate, sand, and carbon is heated in an electric furnace, the phosphate is reduced by the carbon and the element phosphorus obtained. The temperature of the furnace is so high that the phosphorus is volatilized and escapes as a vapor which is led into water and there condensed.

Phosphorus is a pale yellow, brittle solid. It melts at 111° F., and is soluble in carbon disulphide. It oxidizes so readily that it must be kept under water, as it catches fire in the air. It is poisonous, and the constant inhaling of even small quantities of its vapor causes rotting of the bones, and eventually death. It is for this reason that its use in the manufacture of matches has been prohibited in many countries.

When yellow phosphorus is heated to 527° F., out of contact with the air, it changes to a dark red allotropic modification. This red phosphorus is a powder, does not readily oxidize, is insoluble in carbon disulphide, is nonvolatile, and is not poisonous. By heating it to a still higher temperature than that required to form it, it is converted into the usual yellow form. The main use of phosphorus is in the manufacture of matches.

Matches. The head of the ordinary parlor or friction match is composed of a mixture of yellow phosphorus, an oxidizing agent, an abrasive as powdered glass, and glue to hold the mass together. A coloring material is often added, and sometimes sugar to aid the combustion. The top of the wooden splint is sometimes dipped in paraffin to make it inflame more readily. These matches ignite when rubbed on a rough surface because the heat of friction heats the phosphorus to its kindling temperature. The oxidizing agent present aids the combustion, which soon heats the wood of the match stick to its burning point.

These matches are dangerous; for if dropped on the floor and stepped on, they may start a fire. Children, too, sometimes suck the heads of the matches and die from phosphorus poisoning. An improvement is to substitute for phosphorus the less poisonous compound, phosphorus sesquisulphide.

Safety matches have the head composition divided into two parts. Red phosphorus, powdered glass, and glue are on the outside of the match box, while the composition on the head of the wooden splint contains the oxidizing agent, an abrasive, and glue. When these matches are rubbed on the box, the heat developed converts a minute portion of the red phosphorus to yellow, it burns, and the match catches fire. Their great advantage is that they cannot be ignited except on the box. This is not quite true; for if you will carefully rub a safety match on glass, it is possible to heat the head composition to its kindling point, and the match will burn. Under ordinary conditions, however, the composition will wear off by friction before it catches on fire. The wood of the match splint is usually soaked in some chemical that prevents the spark from glowing after the match is extinguished.

Sulphur. Large deposits of sulphur occur in Mexico, Sicily, and Louisiana, as well as in many other parts of the earth. The Louisiana sulphur, which is the source of practically all of the sulphur used in the United States, occurs buried many hundred feet below the surface of the ground. As sulphur is cheap, it would not pay to sink a shaft and mine the sulphur; instead the sulphur is melted in the ground and then pumped out.

To obtain the sulphur, concentric pipes are driven down to the deposit. Through one of these pipes water heated to a temperature of 350° F., under a pressure of 100 pounds to the square inch, is forced. This superheated water melts the sulphur. Through the second pipe, hot compressed air is forced. This forces the melted sulphur, mixed with air bubbles, to come to the surface through the third pipe. When it reaches the surface, it is allowed to flow into large bins inclosed in rough boards, where it solidifies into masses of as much as 100,000 tons. The sulphur produced is over 99 % pure.

Sulphur is a yellow, brittle solid, insoluble in water, but readily soluble in carbon disulphide. Heated in the air it first melts and then burns to sulphur dioxide. It occurs in several allotropic forms, but the two common commercial forms are roll sulphur and flowers of sulphur. Roll sulphur is prepared by melting sulphur and casting it in wooden molds. Flowers of sulphur is prepared by boiling the sulphur and passing the vapor into large brick rooms, where it condenses to a fine powder.

Sulphur is used in enormous quantities in the manufacture of sulphuric acid, in making sulphur dioxide, carbon disulphide, and sulphur dyestuffs, in vulcanizing India rubber, and in gunpowder.

Sulphur dioxide. When sulphur is burned, sulphur dioxide, SO_2, is formed. This is a colorless gas, easily condensed to a liquid. It has a suffocating odor and is a good bleaching and reducing agent. By dissolving it in water sulphurous acid is formed.

$$H_2O + SO_2 \rightarrow H_2SO_3$$

The salts of sulphurous acid, called sulphites, are food preservatives. Sodium sulphite, Na_2SO_3, is a crystalline, white compound, used in almost every photographic developer. Sulphur dioxide is a reducing agent and is widely used as a bleaching and disinfecting agent.

Sulphur trioxide. If a mixture of sulphur dioxide and oxygen is passed over finely divided platinum, they combine and form sulphur trioxide, SO_3. This is a colorless liquid that solidifies at 59° F. If a trace of water vapor is present, the sulphur trioxide solidifies in the form of beautiful silky needles. It must be kept in sealed bottles, as it absorbs water from the air and changes to sulphuric acid.

After the reaction the platinum remains unchanged. It has served a similar purpose to the manganese dioxide used in preparing oxygen; it has hastened a chemical change, but has itself undergone no permanent change. It is a *catalytic agent* or a *catalyzer*. The action is called *catalysis*.

Sulphuric acid. Sulphuric acid is used in so many chemical processes that it may be regarded as the most important acid. It is prepared in two ways, first by the contact and second by the chamber process.

In the contact process sulphur or pyrite is burned to furnish sulphur dioxide. This is mixed with air and passed over iron oxide, which is kept at a temperature of about 650° F. The iron oxide acts as a catalytic agent, converting the sulphur dioxide to trioxide. This is dissolved in dilute sulphuric acid, forming the acid of commerce.

$$S + O_2 \rightarrow SO_2$$
$$SO_2 + O \rightarrow SO_3$$
$$SO_3 + H_2O \rightarrow H_2SO_4$$

The chemical changes involved in the chamber process are very complicated. In general, the change of sulphur dioxide to the trioxide is carried out by the use of oxides of nitrogen as catalytic agents; that is, their alternate oxidation and reduction carry the oxygen to the sulphurous acid, changing it to sulphuric acid.

Sulphuric acid is a colorless, heavy liquid having a specific gravity of 1.84. It boils at 640° F. Its common name is oil of vitriol, for it was formerly made by distilling green vitriol, and the liquid is of an oily appearance.

The action of dilute sulphuric acid upon metals is the same as that of other acids. In hot concentrated solutions, however, it acts as an oxidizing agent. To illustrate: if copper is heated with concentrated sulphuric acid, hydrogen is first set free. This nascent hydrogen is oxidized by the hot acid to water. The sulphuric acid is at the same time reduced to sulphurous acid, which at the high temperature of the reaction breaks up into sulphur dioxide and water.

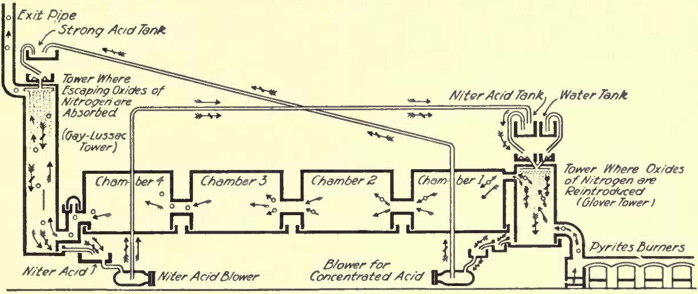

Fig. 126. — Chamber process for the manufacture of sulphuric acid. (From Thorp's *Outlines of Industrial Chemistry.*)

$$\begin{array}{l}
Cu + H_2SO_4 \rightarrow CuSO_4 + H_2 \\
H_2SO_4 + H_2 \rightarrow H_2SO_3 + H_2O \\
H_2SO_3 \rightarrow H_2O + SO_2 \\
\hline
2\,H_2SO_4 + Cu \rightarrow CuSO_4 + 2\,H_2O + SO_2
\end{array}$$

We have seen two illustrations of another use for sulphuric acid, the preparing of other acids. You will recall that you prepared hydrochloric acid by the action of sulphuric acid on sodium chloride, and nitric acid by the action of sulphuric acid upon a nitrate. A similar action occurs whenever sulphuric acid is heated with the salt of another acid, the boiling point of which is lower than that of sulphuric acid; that is, the acid is set free.

Another important use of sulphuric acid is as a drying or dehydrating agent. Gases can be dried by bubbling them through the strong acid, provided of course no chemical change takes place. It would be impossible to dry ammonia in this way, as it would combine with the acid. Dishes containing strong sulphuric acid are often placed in cases containing delicate apparatus to keep the air dry and thus prevent the corrosion of metal parts.

So great is the affinity of sulphuric acid for water that it will withdraw hydrogen and oxygen from organic materials, even where they are not combined in the form of water. A sirup of sugar mixed with strong sulphuric acid will froth up, turn black, and a pasty mass of carbon will be formed. Wood, flesh, fabrics, all are charred in the same way. This makes sulphuric acid dangerous; for if dilute acid is spilled, the water will evaporate, and when the acid becomes concentrated, it will char whatever organic matter it is in contact with.

Sulphates. As sulphuric acid has two replaceable hydrogen atoms, it can form two sulphates, a normal and an acid. Sodium, for example, forms sodium acid sulphate, or sodium bisulphate, $NaHSO_4$, as well as the normal sodium sulphate, Na_2SO_4. The sulphates were formerly called vitriols, and the name still is used with a few compounds. Copper sulphate, $CuSO_4$, is blue vitriol; ferrous sulphate, $FeSO_4$, green vitriol; and zinc sulphate, $ZnSO_4$, white vitriol.

Hydrogen sulphide. Hydrogen sulphide, also called hydrosulphuric acid, H_2S, is prepared by the action of sulphuric acid upon a sulphide. Usually ferrous sulphide is used.

$$FeS + H_2SO_4 \rightarrow FeSO_4 + H_2S$$

It is a colorless gas, about three volumes dissolving in one of water at ordinary temperatures. It has a weak taste, and an exceedingly disagreeable odor, resembling that of rotten eggs. It is a poison, even small quantities producing nausea and headache. It burns, forming sulphur dioxide and water. Chemically it is a weak acid, forming sulphides with metals. Its main use is as a reagent in the chemical laboratory.

Sulphides. The oxidized silver of the jeweler is really silver that has been dipped in a solution of potassium sulphide, and thus a coat of silver sulphide formed on the outside. Many copper articles are finished in the same way. If the coating is polished, it forms a lustrous black, while if it is burnished off in spots, so that the metal shows through, a mottled appearance results. Sodium and potassium sulphides are used in photography to tone bromide prints.

Carbon disulphide. When sulphur vapor is passed over red-hot carbon the two elements combine, forming carbon disulphide. The operation is usually carried out in an electric furnace. In this case, as in most others, the electricity has nothing to do with the chemical change, but is used as a convenient and easily controlled method of heating.

Carbon disulphide is a heavy, colorless liquid, boiling at 115° F. As usually found in commerce it has a nauseating odor, but when pure its odor is pleasant and ethereal. It is very inflammable. It is extensively used as an insecticide, and as a solvent for resins and gums.

Zinc white. A natural mixture of the oxides of iron and zinc called franklinite occurs in New Jersey. If this is heated in a furnace with carbon, the zinc is reduced and volatilized. As this zinc vapor leaves the furnace a current of air is blown into it, when the zinc burns to zinc oxide. The zinc oxide, ZnO, called zinc white, is passed into canvas bags, where the waste gases are filtered out, and the zinc white collected. It is used as a pigment in white paints and as a filler in rubber goods. It has the advantage in paints over lead white that sulphur compounds do not turn it black.

Lead white. Thin perforated discs of lead are placed in small earthenware pots about ten inches high. Dilute acetic acid is then poured in and the pots packed closely together, the space between them and over them being filled with tan bark. The corroding room is filled in this way with row after row of the pots. The process of forming the lead white, $2\ PbCO_3 \cdot Pb(OH)_2$, is a catalytic one. Basic lead acetate first forms; this is decomposed by the carbon dioxide produced by the fermenting tan bark, and basic lead carbonate produced. This sets free the acetic acid, which in turn forms more lead acetate, and so the process goes on. The process requires about three months for its completion. There are quick processes, but the product is thought to be inferior. White lead is the pigment in much paint. It is apt to discolor in city air, as the hydrogen sulphide present converts it into lead sulphide, PbS, a black compound.

The halogens. Fluorine, chlorine, bromine, and iodine form the halogen or chlorine family. Their name comes from a word meaning "I form salts," which is given to them because they all form salts, such as are found in the ocean. Their properties vary in a periodic manner, which follows the order of their atomic weights. The group is an excellent illustration of what is meant by a family in chemistry.

NAME OF ELEMENT	ATOMIC WEIGHT	PHYSICAL STATE	MELTING POINT	COLOR	CHEMICAL ACTIVITY
F	19	gas	– 223°	pale yellow	very very great
Cl	35.5	gas easily liquefied	– 102°	darker yellow green	very great
Br	79.9	liquid	7°	dark red	great
I	126.9	solid	107°	violet black	much less active

If you will examine the above table, you will see that the properties of chlorine are not the same as those of bromine, but that there is a periodic variation in them.

Chlorine is prepared by the oxidation of hydrochloric acid. Bromine then will be prepared in a similar way, but more easily; for since bromine is not as active chemically as chlorine, its compounds will not be as stable.

Bromine is a dark red liquid, soluble in water, and having even a worse effect on the mucous membranes than chlorine. It is used in preparing certain dyes, and in making bromides used in medicine and in photography.

Iodine is a violet black solid, easily volatile, and giving a vapor of a magnificent violet color. It is freely soluble in alcohol, giving the tincture of iodine that we have all used. The iodides are largely used in photography.

Fluorine has no industrial use. The fluorides are of value as a source of hydrofluoric acid, used in etching glass, and calcium fluoride is used as a flux in metallurgy.

Periodic law. If the elements are arranged in the order of their atomic weights, beginning with lithium, a singular fact becomes evident.

Li	Be	B	C	N	O	F	Na	Mg
7	9.1	11	12	14	16	19	23	24.4

We find that the element lithium is a metal. Beryllium is a metal, but is less metallic than lithium; boron is still less metallic, while in carbon the metallic properties are quite lost. As we go on, the elements become more nonmetallic until fluorine is a decided nonmetal. With sodium, however, the pendulum once more swings back, and we have an exceedingly metallic element. If starting with sodium we place it under lithium and go on as before, we find that chlorine comes under fluorine, and then we once more swing back to a metal, potassium.

The vertical rows thus obtained give us the chemical families. The cause of this is not yet known, and the table shows some inconsistencies, yet it has proved of much use to chemists. The above is the barest outline of the facts, but it will give you a foundation that you can build upon by consulting a larger work.

To know the members of the chemical families in the order of their atomic weights is useful because it helps us to remember their properties. To illustrate: if I know that sodium and potassium are both members of the same family and follow each other in that order, I then know that potassium will be more active than sodium, that potassium hydroxide will be a stronger base than sodium hydroxide, etc. Note that the properties of the members of a family are not the same, but that they show a regular gradation as we go from element to element in the same family, and that this applies to both physical and chemical properties. It is

PERIODIC ARRANGEMENT OF THE ELEMENTS

GROUP	0	I		II		III		IV		V		VI		VII		VIII		
Highest Oxide	E	E_2O		EO		E_2O_3		EO_2		E_2O_5		EO_3		E_2O_7		EO_4		
Family		A	B	A	B	A	B	A	B	A	B	A	B	A	B			
First Short Period	**He** (4)	**Li** (7)		**Be** (9.1)		**B** (11)		**C** (12)			**N** (14)		**O** (16)		**F** (19)			
Second Short Period	**Ne** (20)	**Na** (23)		**Mg** (24.3)		**Al** (27.1)		**Si** (28.3)			**P** (31)		**S** (32.1)		**Cl** (35.5)			
First Long Period	**A** (39.9)	**K** (39.1)	**Cu** (63.6)	**Ca** (40.1)	**Zn** (65.4)	**Sc** (44.1)	**Ga** (69.9)	**Ti** (48.1)	**Ge** (72.5)	**V** (51.2)	**As** (75)	**Cr** (52)	**Se** (79.2)	**Mn** (54.9)	**Br** (79.9)	**Fe** (55.9)	**Co** (59.0)	**Ni** (58.7)
Second Long Period	**Kr** (83)	**Rb** (85.5)	**Ag** (107.9)	**Sr** (87.6)	**Cd** (112.4)	**Yt** (89)	**In** (114.8)	**Zr** (90.6)	**Sn** (119)	**Cb** (93.5)	**Sb** (120.2)	**Mo** (96)	**Te** (127.5)	——	**I** (126.9)	**Ru** (101.7)	**Rh** (102.9)	**Pd** (106.7)
Third Long Period	**Xe** (130.2)	**Cs** (132.8)	**Au** (197.2)	**Ba** (137.4)	**Hg** (200)	**La** (139)	**Tl** (204)	**Ce** * (140.3)	**Pb** (207.1)	**Ta** (181)	**Bi** (208)	**W** (184)	——	——	——	**Os** (190.9)	**Ir** (193.1)	**Pt** (195)
	Nt (222.4)			**Ra** (226.4)				**Th** (232.4)				**U** (238.5)						

*Between Cerium and Tantalum lie a number of elements which are very similar to each other in properties and which, together with **Yttrium, Lanthanum, and Cerium,** which they much resemble, constitute the so-called rare earth elements.

difficult to give credit to the author of such a generalization as this, for many men have worked on it, each improving the idea a little. The Russian chemist, Mendeléeff, was, however, the first who published such a table (1869).

Milk an important food. The fact that infants thrive on an exclusive diet of milk shows that it must contain all of the elements needed by the body; that is, that it must be a complete food. It is so largely used by young children, and is so easily contaminated with germs, and so difficult to keep, that great care should be used in selecting your milkman.

Many cities have adopted rigid rules for the regulation of the sale of milk, for it has been conclusively shown that children's diseases and the sale of poor milk go hand in hand. In New York there are three grades, A, B, and C, sold. Your milkman can tell you the local regulations, and usually they are printed by the Board of Health for free distribution. Read them and see if you know the reason for the various restrictions imposed on the dealers.

Composition of milk. In various parts of the world the milk of goats, mares, and reindeer is used, but to us milk always means cow's milk. Its average composition is water 87.17, sugar 4.88, fat 3.69, protein 3.55, and mineral matter 0.71. The composition varies within wide limits, depending on the time of year and the breed of the cow, as well as on the individual cow from which the milk is drawn. In general the laws will not allow a milk to be sold that contains less than 3% of butter fat, and 11.5% of total solids. Some cows have given over 20,000 pounds of milk a year, but the average is less than one fourth of this.

Preserving. Milk is an ideal culture medium for bacteria. To preserve it, it must be carefully protected from every source of contamination, and kept cold. The addition of preservatives, as formaldehyde and boric acid, is very objectionable and is prohibited by law.

Pasteurized milk. To sterilize milk it must be heated to the boiling point, allowed to stand and then heated again to the boiling point. This cooks the milk, and so alters its taste that the process is impracticable. If, however, the milk is heated to 167° F., for 20 minutes, and then rapidly cooled, the taste is not perceptibly altered, and the lactic acid bacteria are killed. The milk will then keep sweet for several days, if it is kept cold. Pasteurized milk is not sterile, and the putrefactive bacteria present will multiply, even though the milk does not turn sour.

Condensed milk. Much of the milk produced is turned into condensed milk. The milk is placed in vacuum pans, and water evaporated until a thin liquid containing about 28 % of milk solids remains. This is canned, heated again to sterilize it, when it keeps for a long time. Sugar is often added, when the sugar will prevent fermentation.

Milk powders. Many attempts have been made to evaporate milk and sell it in the form of a powder. These all failed for a time owing to the difficulty of evaporating such a mixture to dryness, and reproducing the taste of raw milk in the redissolved powder. Recently the problem has been solved by blowing milk through fine holes, so as to produce a very fine spray, into a hot room. The hot air evaporates the water from each tiny drop, and the milk powder remains. It is freely soluble in water, and gives a product that while not identical with the raw milk, is close to it, and is perfectly satisfactory for cooking. The powder keeps better if skimmed milk is used, as the butter fat causes the product to become rancid. Its usefulness in camping, or to sailors, etc., is evident.

Homogenized milk is milk that has been forced through holes much smaller than the average diameter of the fat globules present in milk. A pressure of about 2500 pounds to the square inch is necessary to do this. The result is that the fat globules are made much smaller, and a more perfect emulsion obtained, that does not separate as does milk. That is, the cream does not readily rise. The milk is much thickened by the process. The homogenized milk is largely used in the manufacture of ice cream.

Butter. Cream that has ripened, *i.e.* slightly soured, when churned gives butter. The churning process brings the fat globules together, leaving the other constituents of the milk in the buttermilk, which has about the same composition as skimmed milk. The butter is worked to free it from water, salted, packed, and sold.

Cheese. Rennet, which is an extract from the fourth stomach of the calf, obtained by soaking the stomach in a dilute salt solution, has the power, owing to a ferment that it contains, of causing the casein of milk to turn solid. Junket tablets contain rennet, and cause this change. This is the first step in the making of cheese.

Rennet is added to milk, and after the mass has set and the curd has become firm, it is cut into cubes and allowed to stand. The curd shrinks, becomes harder, and the whey separates. The curd is then piled, when it forms a solid mass, more whey draining from it. It is then ground, salted, molded, stored until the ripening process develops a satisfactory flavor, and marketed.

The countless varieties are due to a slight extent to the different milks used, but mainly to the ripening process, different molds and bacteria giving different flavors.

Fermented milk. Milk that has soured has an acid taste, agreeable to many. Buttermilk, which is sour skimmed milk, has a large sale. Kumiss, Fermalac, and Zoolak are artificially soured milks. They are said to be easily digestible.

Oxalic acid. Potassium and calcium oxalates are found in plants, as the rhubarb and sorrel. The acid is prepared by heating a thick paste of sawdust and sodium hydroxide on iron plates. Sodium oxalate is formed, from which the acid is extracted.

Oxalic acid, $H_2C_2O_4$, crystallizes with two molecules of water. It is a poison. As it freely dissolves the oxides of metals, it is much used in the home in metal polishes for brass and copper.

Tartaric acid. Many plants contain the acid, $H_2C_4H_4O_6$, tartaric acid. Commercially it is obtained from the impure potassium hydrogen tartrate (argols) that separates during the fermentation of wine. The acid is used in dyeing, and some of its salts are important. Seidlitz powders contain potassium acid tartrate in one paper, and sodium acid carbonate in the other. On mixing, Rochelle salt, carbon dioxide, and water are formed. With the use of tartrates in baking powders we are familiar.

Boracic or boric acid. Boracic acid, H_3BO_3, is a feeble acid. Its solution is used as an eye wash, and as a food preservative.

Tannic acid. The astringent principle found in oak and hemlock bark and in many plants is called tannin. It is a mixture of a number of compounds, and is used in dyeing, making writing ink, and tanning leather.

Nutgalls contain much tannin. If you boil a small quantity of powdered nutgalls in water, and add a little ferrous sulphate, a violet-colored precipitate, slowly changing to black, forms. When thickened by the addition of a little gum, this forms one of the common inks. The ink would be too pale when first used, so a color is usually added to it. As the iron tannate gradually forms, the ink turns black on the paper.

Leather is made by first cleaning and dehairing the hide, and then soaking it in a tannin solution. The gelatinous compounds of the hide are gradually made insoluble by combining with the tannin, and leather results. There are other methods of making the hide insoluble, as by the use of chromium compounds.

SUMMARY

Aluminium oxide is the mineral corundum. Ruby and sapphire are impure forms used as gems. Emery is an impure form used as an abrasive. Alundum is fused aluminium oxide and is used as an abrasive.

An abrasive is a material used to grind and polish hard substances.

Carborundum is silicon carbide and is one of the best abrasives.

Thermit is a mixture of aluminium and iron oxide. It gives on ignition an intense heat.

An alum is a double sulphate of aluminium and some other metal.

Nitric acid is made by heating sulphuric acid and a nitrate together. Nitrates are good oxidizing agents.

Nitroglycerin is glyceryl nitrate. It is a high explosive.

Dynamite is nitroglycerin absorbed in some material, as ground wood.

Guncotton is cotton hexanitrate.

Phosphorus is made by heating calcium phosphate and carbon in an electric furnace. Red and yellow are two allotropic forms.

Matches are phosphorus or phosphorus sesquisulphide, an oxidizing material, glue, and an abrasive. Safety matches have the composition divided into two parts, and one part is placed on the match box.

Sulphur is obtained from Louisiana by melting the deposit in the ground and forcing it to the surface with compressed air. It is a yellow solid, insoluble in water but soluble in carbon disulphide. Used in making sulphuric acid, vulcanizing rubber, and in dyestuffs.

Sulphur dioxide is prepared by burning sulphur. It is used in bleaching and disinfecting.

Sulphuric acid is made by oxidizing sulphur dioxide and dissolving sulphur trioxide produced in water. It is the most important acid.

Carbon disulphide is prepared by the direct union of carbon and sulphur. It is used as a solvent and insecticide.

Zinc white is zinc oxide and is used as a pigment.

Lead white is basic lead carbonate and is used as a pigment.

Hydrogen sulphide is used in the laboratory as a reagent.

Oxidized silver is silver coated with silver sulphide. Other metals are colored in a similar way.

Milk is a universal food, easily spoiled and must be preserved with great care.

Pasteurizing kills the lactic acid bacteria, and preserves milk. It does not sterilize the milk.

Condensed milk is milk from which some of the water has been evaporated.

Milk powder is milk, usually skimmed, from which all of the water has been taken.

Homogenized milk is milk that has been forced through minute orifices to make it a more perfect emulsion.

Cheese is the dried and ripened curd of milk, produced by adding rennet to milk.

Butter is the collected butter fat of milk.

Modified milk is milk so treated as to make its composition resemble that of mothers' milk.

Oxalic acid is a poison. It is used in metal polishes.

Tartaric acid is made from argols. Tartrates are used in baking powders and dyeing.

Boracic acid is a food preservative.

Tannin is used in making leather and ink.

Leather is hide made insoluble by tannin.

Exercises

1. Why is sulphur not mined instead of being obtained by the present method?

2. How could you prepare flowers of sulphur in the laboratory?

3. How could you blacken a copper ash tray?

4. Which is the more harmful acid if spilled upon a coat, dilute sulphuric or hydrochloric acid? Why?

5. What compound gives rotten eggs their disagreeable odor? Account for the formation of this compound.

6. How would you fumigate a coat, using sulphur dioxide?

7. Sulphur dioxide is sometimes used in drying fruits. Do you consider this objectionable?

8. How could you obtain sulphur dioxide from sulphuric acid in the laboratory?

9. Have you any reason for thinking that a strong tincture of iodine might be injurious? Explain.

10. As sulphuric acid takes no part in the chemical change, why is it used in the preparation of nitroglycerin?

11. Carbon disulphide is used to kill mice in grain. Is there any danger in using it? What precautions would you take?

12. Why does ammonia not restore the color of a dress stained with nitric acid?

13. Would you obtain hydrogen by the action of strong nitric acid upon copper? Explain.

14. Why is the paper label on the outside of the sulphuric acid bottle often blackened?

15. Give two reasons for the severe burns caused by sulphuric acid.

16. Cæsium, atomic weight 132.8, is a member of the sodium family. Tell its properties.

17. Rubidium, atomic weight 85.5, is a member of the sodium family. Will it decompose water more or less energetically than potassium? How do you know?

18. Why is milk often the transmitting agent for typhoid fever?

19. If formaldehyde will preserve milk, why not use it?

20. Is there any difference between pasteurized and sterilized milk? Explain.

21. How would you obtain milk sugar, starting with raw milk?

APPENDIX

METRIC MEASUREMENTS

The nomenclature in the metric system is extremely simple, since there is one general principle which applies to all tables of measurements. The Greek prefixes for 10, 100, 1000, viz., *deca, hecto,* and *kilo,* are used to signify multiplication; while the Latin prefixes for 10, 100, and 1000, viz., *deci, centi,* and *milli,* are employed to express division.

The principle is illustrated in the following table of linear measure:

LINEAR MEASURE

Divisions . .	MILLIMETER	=	0.001	meter,	$\frac{1}{1000}$	of a meter
	CENTIMETER	=	0.01	"	$\frac{1}{100}$	"
	Decimeter	=	0.1	"	$\frac{1}{10}$	"
Unit . . .	METER	=	1	"		
Multiples . .	Decameter	=	10	"		
	Hectometer	=	100	"		
	KILOMETER	=	1000	"		

The only units of cubic capacity that we need to use are the *cubic centimeter* (cc.) and the *liter* (l.) which is 1000 cc. and is roughly equivalent to 1 quart.

The weight of 1 cc. of distilled water at 4° C. is called 1 *gram.* A *kilogram* is 1000 grams. 1 liter of water weighs 1 kilogram.

METRIC AND ENGLISH EQUIVALENTS

1 m.	= 39.37 in.	1 in.	= 2.5400 cm.
1 m.	= 3.2809 ft.	1 ft.	= 30.480 cm.
1 m.	= 1.0936 yd.	1 yd.	= 0.91440 m.
1 Km.	= 0.62137 mi.	1 mi.	= 1.60935 Km.
1 sq. cm.	= 0.1550 sq. in.	1 sq. in.	= 6.452 sq. cm.
1 sq. m.	= 10.764 sq. ft.	1 sq. ft.	= 929.0 sq. cm.
1 sq. m.	= 1.196 sq. yd.	1 sq. yd.	= 0.8361 sq. m.
1 cc.	= 0.06102 cu. in.	1 cu. in.	= 16.387 cc.
1 cu. m.	= 35.314 cu. ft.	1 cu. ft.	= 28316 cc.
1 cu. m.	= 1.3080 cu. yd.	1 fluid oz.	= 29.574 cc.
1 l.	= 0.26417 gal.	1 liquid qt.	= 0.94636 l.
1 l.	= 0.9081 dry qt.	1 dry qt.	= 1.101 l.
1 g.	= 15.4324 gr.	1 gr.	= 0.0648 g.
1 Kg.	= 2.2046 lb. avoir.	1 lb. avoir.	= 453.59 g.
1 g.	= 0.03527 oz.	1 oz. avoir.	= 28.3495 g.

PHYSICAL CONSTANTS OF COMMON ELEMENTS

Element	Symbol	Atomic Weights		Valence	Specific Gravity		Melting Point	Boiling Point
		Approximate	Exact O = 16		Water=1	Air=1	° C.	° C.
Aluminium	Al	27	27.1	III	2.6		660	1600
Antimony	Sb	120	120.2	III V	6.6		630	1600
Argon	A	40	39.88	——		1.38	−188	−186
Arsenic	As	75	74.96	III V	5.7		185	449.5
Barium	Ba	137	137.37	II	3.8		850	950
Bismuth	Bi	208	208.0	III V	9.7		269	1435
Boron	B	11	11.0	III	2.6		infusible	3500
Bromine	Br	80	79.92	I	3.1		−7.3	63
Cadmium	Cd	112	112.4	II	8.6		322	778
Calcium	Ca	40	40.07	II	1.5		800	——
Carbon	C	12	12.00	IV	1.7–2.1		sublimes	3500
Chlorine	Cl	35.5	35.45	I		2.49	−102	−33.6
Chromium	Cr	52	52.0	II III VI	6.9		1515	——
Copper	Cu	63.6	63.57	I II	8.7		1090	2100
Fluorine	F	19	19.0	I		1.31	−223	−187
Gold	Au	197	197.2	I III	19.3		1064	——
Helium	He	4	4.0	——		0.13	−271	−267
Hydrogen	H	1	1.008	I		0.07	−256.5	−252.5
Iodine	I	127	126.92	I	4.9		114	184
Iron	Fe	56	55.84	II III	7.8		1804	——
Lead	Pb	207	207.1	II IV	11.3		327	1500
Magnesium	Mg	24.3	24.32	II	1.7		632	1100
Manganese	Mn	55	55.0	II IV	7.4		1245	——
Mercury	Hg	200	200.6	I II	13.6		39.5	357
Nickel	Ni	58.7	58.68	II	8.7		1484	——
Nitrogen	N	14	14.01	III V		0.96	−210	−196
Oxygen	O	16	16.00	II		1.10	−230	−182
Phosphorus	P	31	31.04	III V	1.8		44.2	290
Platinum	Pt	195	195.2	IV	21.5		1750	——
Potassium	K	39	39.10	I	0.87		62.5	757
Silicon	Si	28	28.3	IV	2.0		1400	3500
Silver	Ag	108	107.88	I	10.5		961	2050
Sodium	Na	23	23.	I	0.97		97.6	877
Sulphur	S	32	32.07	II IV VI	2.0		114.5	445
Tin	Sn	119	119.0	II IV	7.3		232	1500
Zinc	Zn	65.4	65.37	II	7.1		419	918

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