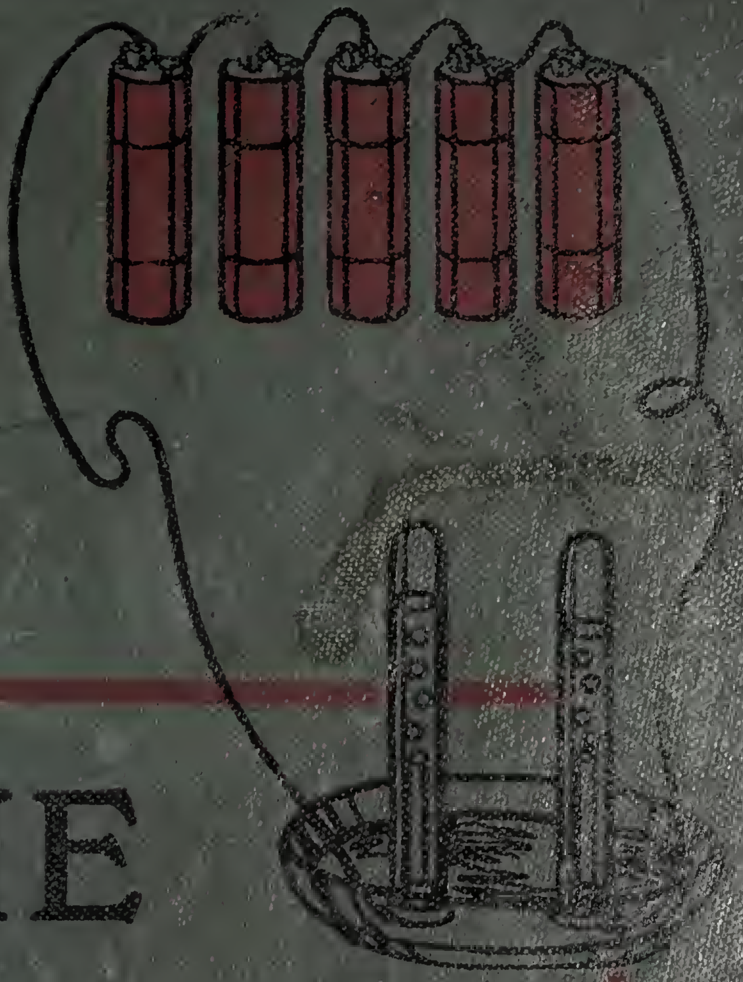
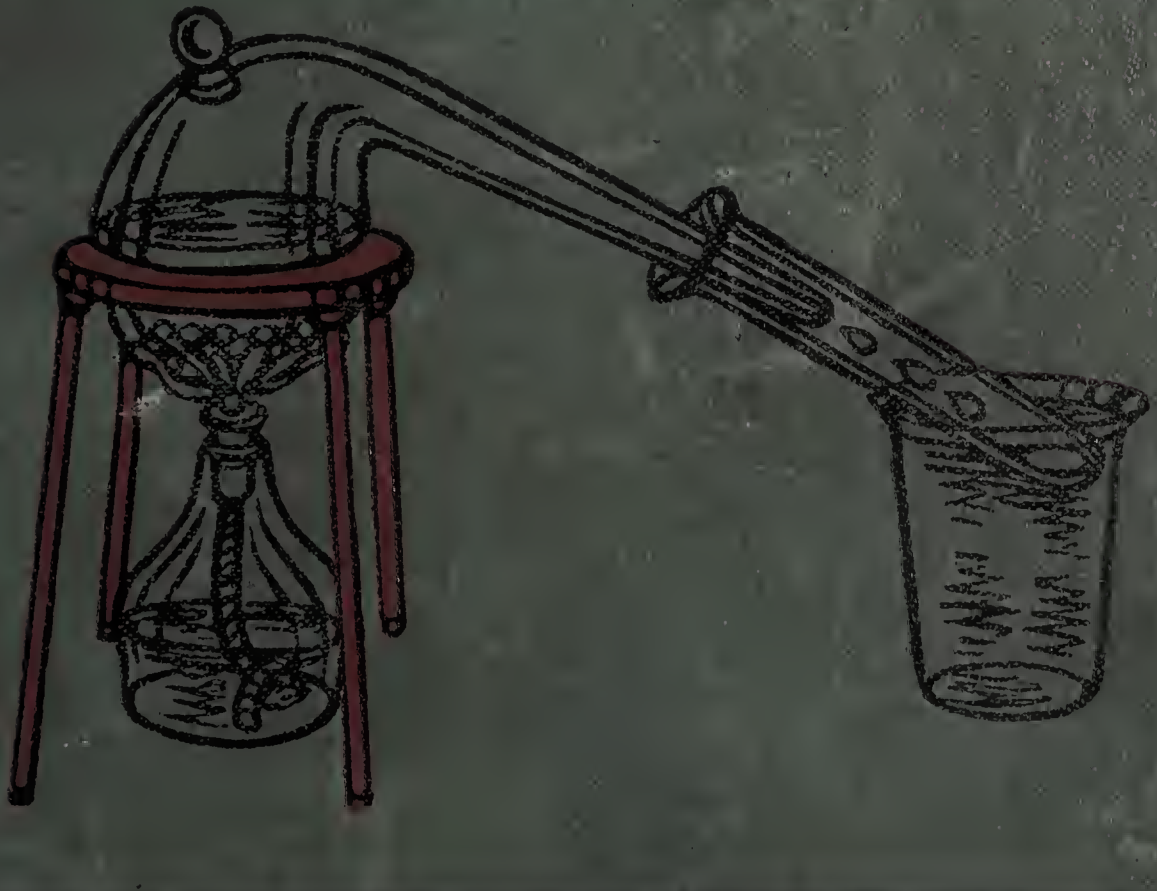


THE BOY CHEMIST

A. FREDERICK COLLINS

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THE BOY CHEMIST

The Alchemist.

(From Painting by Teniers.)

THE BOY CHEMIST

BY

A. FREDERICK COLLINS

Author of "The Boy Astronomer."

WITH FRONTISPIECE AND ONE HUNDRED AND SEVENTY-TWO TEXT ILLUSTRATIONS

1830 1904 1861
KNOWLEDGE NO MORE SHALL BE A FOUNTAIN CLOSED

LOTHROP, LEE & SHEPARD CO.
BOSTON

THE BOY CHEMIST

Printed in the U.S.A.

TO

WILLIAM COX

FOREWORD

During the dark ages there lived men who worked over seething furnaces in the vain attempt to find a way to live a thousand years and to make precious metals out of the baser ones. These men were the *alchemists* of old, and as they lived in an age of ignorance they believed many strange and untrue things.

Chief among these were that *earth*, *air*, *fire*, and *water* were elements; that when these acted one on another, *sulphur*, *mercury*, and *salt* were formed, which in turn were called *principles;* and that by combining the elements and the principles in the right proportions they would yield an *essence*, which when taken internally, like castor oil, would prolong life indefinitely, and when poured over lead would change, or "transmute," it into pure gold.

Of course they did not find this miraculous essence, so diligently sought, but they made a few simple discoveries which, while seeming to them to be of very little value, were really of priceless worth, for out of them came the great science of Chemistry as we know it to-day. There is nothing in the whole realm of knowledge of more absorbing interest in an experimental way than this branch of it, for it tells you how to combine atoms and molecules of various elements produced by Nature and make other and entirely different substances. Many of these substances Nature herself has never made, so that you have it in your power to be a creator in the smaller sense; and, equally wonderful, Chemistry tells you how to *break down* various compounds by means of light, heat, and electricity, and separate them into their original elements.

If you would like to do these things, you will like Chemistry, and I have written this book so that you can easily make the experiments, and thus gain a very good idea of how and why chemicals react on each other, especially if you will give careful attention to the chapter "Chemistry Simply Explained," and do a little *thinking* at the same time. You do not need to buy all the apparatus and chemicals at once, but just enough to make a few of the simple experiments that interest you at first. The next step is to fit up a little laboratory, put on an apron, and go to work in earnest. If you make the experiments in order, by the time you reach the tenth chapter you will have taken a fairly good course in Chemistry, and one that will serve you well for all time.

From the tenth chapter on, you will find a large number of experiments that are strange and curious in the extreme, such as making a flame without light, making photographs with a pinhole camera, magical experiments of all kinds, and making safe and sane fireworks. Finally, there are numerous experiments that have to do with household chemistry, which are of much interest and have a great deal of practical value. That much pleasure and profit will accrue to you from this book is the wish of the author.

A. FREDERICK COLLINS.

Paris, France.
May, 1924.

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ILLUSTRATIONS

THE BOY CHEMIST

CHAPTER I

WHAT YOU NEED TO EXPERIMENT WITH

The two chief things you need to make the experiments described in this book are *the apparatus* and *the chemicals.* You can improvise some of the apparatus and use household china and glassware for other pieces, but it is better to buy whatever equipment you need, for it costs but very little, and having been designed especially for the purpose, it will prove far more satisfactory. However, I shall tell you how to construct whatever pieces of apparatus I think you can make, so that in case you happen to be far removed from a chemical supply house you can go ahead with the experiments, anyway, and not lose valuable time.

The Apparatus You Need. If you can, it is the better way to get the following pieces of apparatus before you start in to experiment, as you will need them right along:

1. A Ring-Stand.
2. An Alcohol Lamp, or a Bunsen Burner.
3. A Sheet of Iron Gauze.
4. A Set of 6 Test Tubes.
5. A Test-Tube Rack.
6. A Test-Tube Brush.
7. A Glass Stirring Rod.
8. A Pipette, or Medicine Dropper.
9. An Ordinary Teaspoon and a Tablespoon.
10. A Nest of 3 Beakers.

11. One or Two Flasks.
12. A Glass Funnel.
13. A Mortar and Pestle.
14. Two or Three Wide-Mouth Bottles.
15. Some Corks, or, better, Rubber Stoppers to fit the Bottles.
16. A Cork-Borer.
17. A Graduated Glass.
18. A Porcelain Crucible.
19. A Glass Retort.
20. Several Watch Glasses.
21. A Glass U-Tube.
22. Several Pieces of ⅜-Inch Glass Tubing.
23. Some ¼-Inch Rubber Tubing.
24. A Pair of Forceps, or Tweezers.
25. A Dozen Sheets of Filter Paper.
26. Two Sheets of Litmus Paper.

What the Apparatus Consists of. Now before we go any farther, let us find out just what each of the above pieces of apparatus consists of.

A bought ring-stand is shown in Fig. 1, and one that you can make in Fig. 2. Take a piece of ⅛-inch or 3/16-inch iron wire 2 feet long, and form a ring on one end 4 inches in diameter so that it will stand on the table. Now take a piece of ⅛-inch iron wire and make a ring 2¾ inches in diameter and then bend the free end into a spiral of three or four turns with a pair of round-nose pliers, so that it will slip snugly over the support rod of the ring-stand, and it is ready to use.

The proper kind of an alcohol lamp to use is shown in Fig. 3, but if you are hard-pressed for a flame you can make a lamp of an empty inkstand; to do so, bend a strip of tin 1¼ inches wide and 1½ inches long into a tube and then bore a 5/16-inch hole through the cork and push the tube into it so that it will project at both ends. Next, make a wick of string about 4 inches long and put this in the tube, half fill the bottle with methyl alcohol (CH_3OH), which is ordinary *wood alcohol*, put the long end of the wick into it, and then the cork in the bottle, and your alcohol lamp will look like Fig. 4.

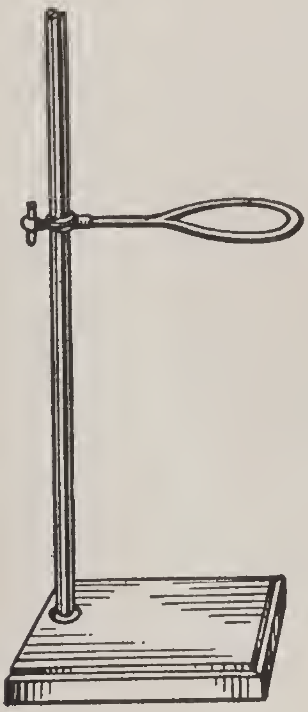

FIG. 1.—A Bought Ring-Stand.

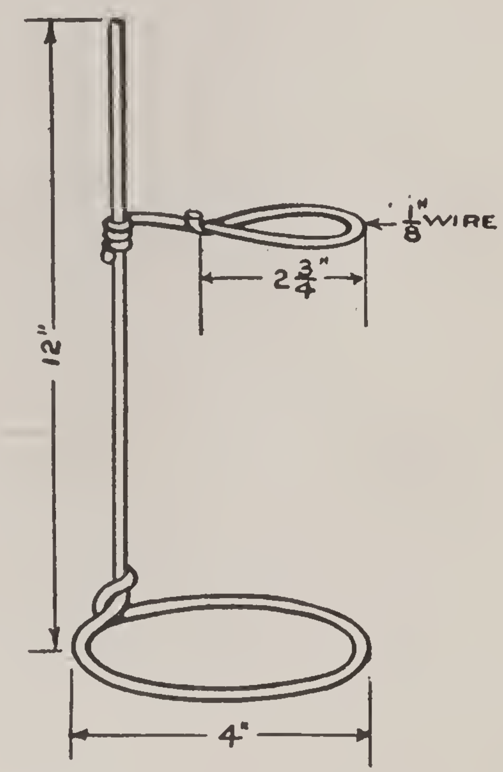

FIG. 2.—A Home-Made Ring-Stand.

You can buy a Bunsen burner, Fig. 5, for 50 cents or less, and this will give you a much hotter flame than an alcohol lamp. Again, if you are pressed for the want of one, take a piece of iron pipe ⅜ inch in diameter, inside measurement, and 6 inches long, and drill a ¼-inch hole through it about 1½ inches from one end. Then make a tin ring ½ inch wide that will just slip over the pipe and cover the hole, and you can regulate the supply of air.[1]

Now bore a hole in the center of a block of wood that is ½ inch thick and 3 inches on the sides and push the tube into it; glue two strips of wood, ½ inch thick, to the bottom of the block along its edges, then put a rubber tube on the lower end of the pipe and connect it with a gas jet from which you have taken out the tip, and the burner is complete, as shown in Fig. 6. A sheet of iron gauze is laid on the upper ring of the stand and the flask, or other piece of chemical glassware, is set on it when you want to heat any liquid to the boiling point. Iron gauze comes in

FIG. 3.—A Bought Alcohol Lamp.

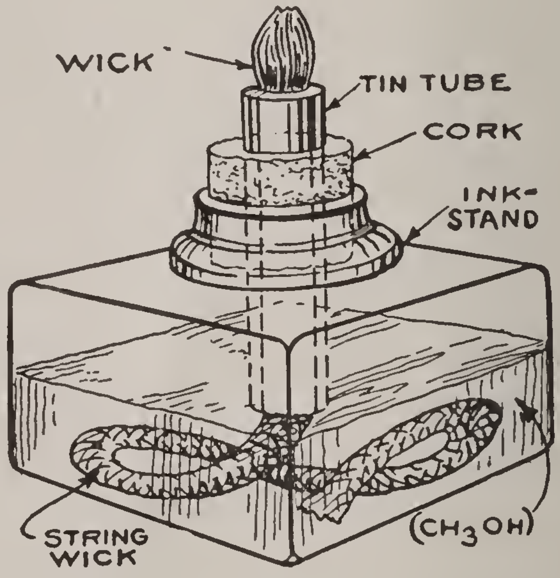

FIG. 4.—A Home-Made Alcohol Lamp.

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[1] The reason for regulating the air supply is explained in Chapter XI.

sheets 4 by 4 inches on the sides, and you can also buy it with the meshes filled in with asbestos for slow evaporation.

A set of three 5-inch and three 6-inch test tubes, see Fig. 7, will serve for your experiments, at least at first. To clean the test tubes you will need a test-tube brush, as shown in Fig. 8, and you can buy or make a test-tube holder, as shown in Fig. 9. A regular test-tube rack is shown in Fig. 10, but you can make one by bending a piece of $^3/_{32}$-inch brass or iron wire 4 feet long into the shape shown in Fig. 11.

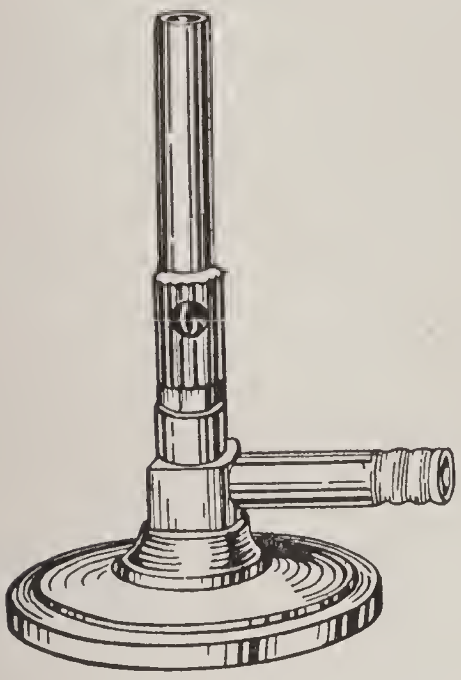

Fig. 5.—A Bought Bunsen Burner.

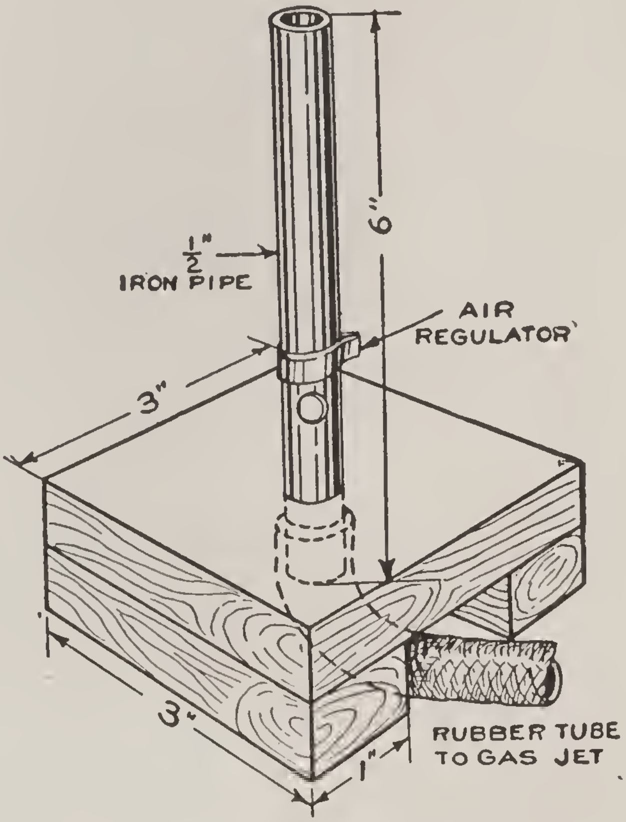

Fig. 6.—A Home-Made Bunsen Burner.

A glass stirring rod is a solid glass rod about $^3/_{16}$ inch in diameter and 6 inches long, as shown in Fig. 12. You can use a glass tube instead, though some of the liquid usually gets up into it and stays there. A pipette is simply a *medicine dropper*, see Fig. 13, and this enables you to put one or more drops of a liquid into a test tube or beaker with neatness, accuracy, and dispatch.

A beaker is a tumbler-shaped glass with a lip on it, as shown in Fig. 14, so that you can pour a liquid from it without spilling it. You can use an ordinary glass tumbler instead of a beaker except when you have to heat it. You should have a set of three beakers, or *nest*, as it is called, because one goes inside another, and the smallest can be 2 inches, the next 2½ inches, and the third 3 inches in diameter.

FIG. 7.
A Test Tube.

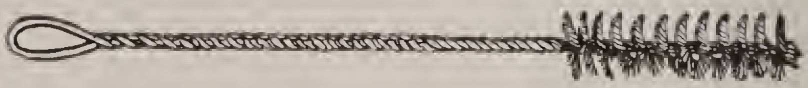

FIG. 8.—A Test-Tube Brush.

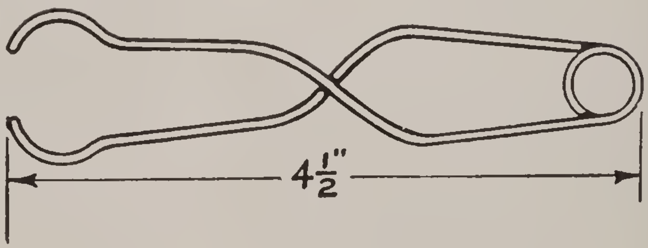

FIG. 9.—A Test-Tube Holder.

You can buy in two shapes flasks of annealed glass that can be heated without breaking, and these are shown in Figs. 15 and 16. The first is the regular spherical form with a flat bottom, and this can be set directly in the ring of your stand over the flame. The second is called an *Erlenmeyer* flask, and its shape is such that while it cannot be easily tipped over, it can be set only on a ring-stand on a piece of wire gauze.

A glass funnel will be found useful where you have to transfer the contents of one vessel into another, especially if these are bottles, as well as for filtering solutions. A funnel with a mouth 3 inches in diameter, and with the stem cut at an angle, as shown in Fig. 17, will be large enough. You will need also a package of 5-inch filter paper. A 2½-inch or 3-inch glass or porcelain mortar and

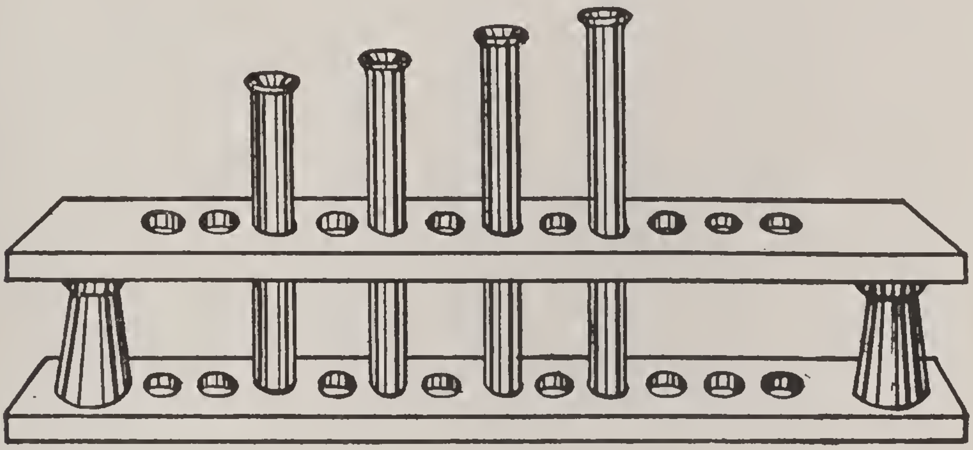

FIG. 10.—A Bought Test-Tube Rack.

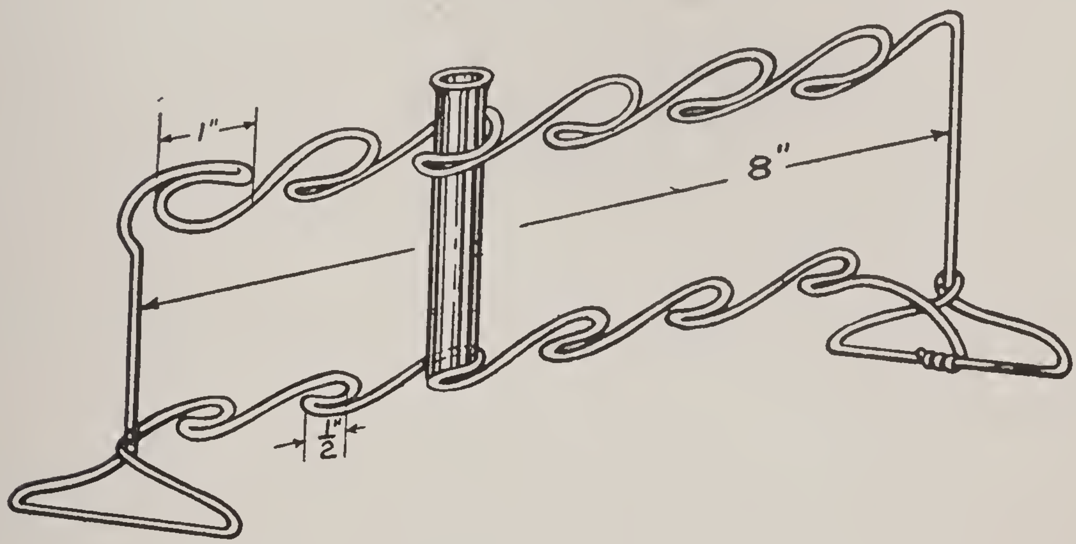

FIG. 11.—A Home-Made Test-Tube Rack.

a pestle, see Fig. 18, must be used where you have to grind a substance to a powder.

A 4-ounce or an 8-ounce wide-mouth bottle, like that pictured in Fig. 19, is used in many operations, especially in purifying gases. You can use a ½-pint fruit jar in a pinch, but a bottle is better. An ordinary cork will serve as a stopper, but a rubber stopper makes a tighter fit. While you can make a hole in a cork with a knife and smooth it up with a rat-tail file, an easier, quicker, and, hence, better way is to use a cork-borer. Rubber stoppers can be bought with holes in them, ready for inserting glass tubes. In experiments where two or three glass tubes have to be inserted in a bottle, you can also use a *Woulff's bottle* that has two or three necks, as shown in Fig. 20.

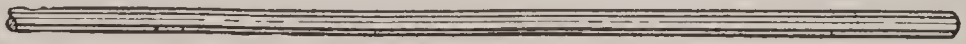

FIG. 12.—A Glass Stirring Rod.

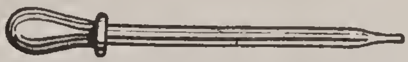

FIG. 13.—A Pipette or Medicine Dropper.

With a graduated glass, Fig. 21, you can measure out liquids in fluid ounces, and you can get at the drug store a small one with which you can measure 1 to 8 teaspoonfuls and from 1 to 2 tablespoonfuls; it is shown in Fig. 22, and for many of the experiments described in this book it is very convenient. A porcelain crucible fitted with a cover enables you to heat compounds to a high temperature; a small one having a diameter of 1¼ inches, see

Fig. 23, will be quite large enough for any experiment you will want to make.

The retort shown in Fig. 24 is made of a kind of glass which has been carefully annealed, so that it can be heated to quite a high temperature without breaking. It is made with a ground-glass stopper, and can be set in a ring-stand directly over a flame. Half a dozen watch glasses, or *crystals*, as they are often called, see Fig. 25, are useful for evaporating small quantities of solutions, and you should also have a 3-inch porcelain evaporating dish, which is shown in Fig. 26. A 5-inch U-tube, pictured in Fig. 27, is a convenient apparatus for purifying and drying gases. When handling phosphorus (*P*), potassium (*K*), and other solids that you do not want to get on your hands, use a pair of forceps, or *tweezers*, as they are commonly called. These are shown in Fig. 28.

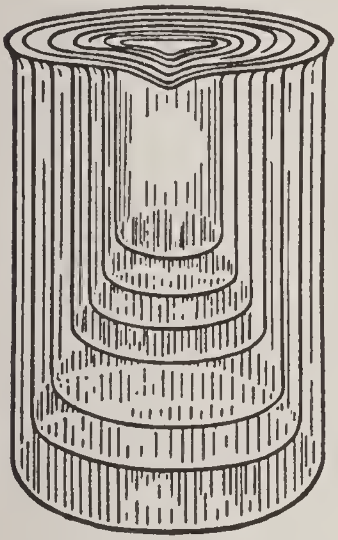

Fig. 14.—A Nest of Beakers.

Fig. 15.—An Ordinary Spherical Flask.

Fig. 16.—An Erlenmeyer Flask.

Besides the above apparatus you will need several pieces of glass tubing that is sold under the trade name of *German soft glass;* you can easily bend this kind of glass in the flame of your alcohol lamp or Bunsen burner, which process I shall explain to you a little farther along. Get the tubing ⅜ inch (about 8 millimeters)[1] in diameter, outside measurement, and in 2-foot lengths. You can buy glass T-tubes and glass Y-tubes of the same size if you should need them. You must also get two or three feet of rubber tubing, with inside diameter, ¼ inch, for making connections.

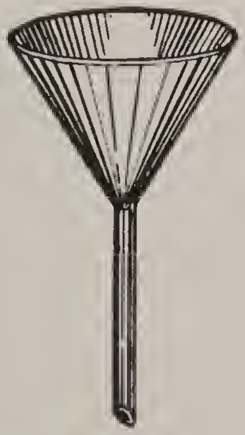

FIG. 17.—A Glass Funnel.

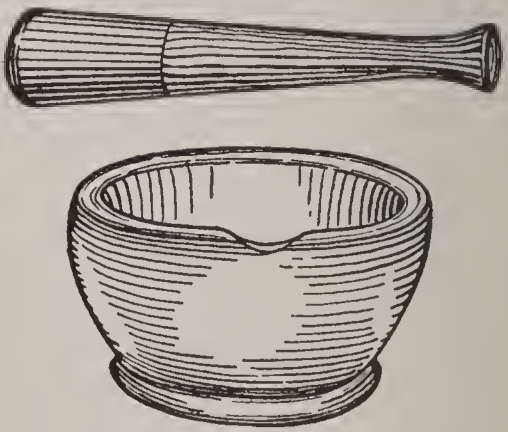

FIG. 18.—A Mortar and Pestle.

While I have given the amounts of chemicals to be used in making the experiments in this book in the way that seemed to me to be the easiest for you to measure out, still, in chemistry as it is taught in schools to-day, the solid kinds of chemicals are most carefully weighed out on a pair of scales, or a balance, and liquids are measured out in a graduated cylinder, and for both of these the *metric system* of measurements is used.

This system of weights and measures runs in multiples

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[1] It is listed in the catalogues in the *metric system.*

of ten, and this is, consequently, much more simple than the English system, which is purely an arbitrary one that has come down to us from an unscientific past. You can buy a small pair of hand scales, or a balance, for $1.50 or so, but the weights will cost you considerably more, and you can get a medium-size graduated cylinder for about $.75. Both of these you will eventually require in order to do accurate work, but for all the experiments that follow in these pages they will not be needed.

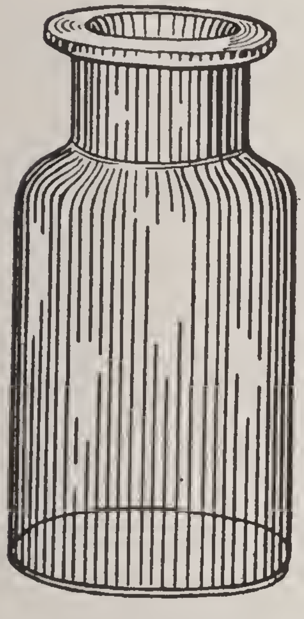
FIG. 19.—An Ordinary Wide-Mouth Bottle.

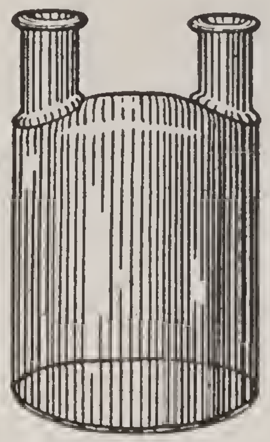
FIG. 20.—A Woulff's Bottle.

FIG. 21.—A Four-Ounce Graduated Glass.

FIG. 22.—A Teaspoonful Graduated Glass.

Your Supply of Chemicals. I shall not write out a list of the chemicals you will need, for you can do this better after you have decided what experiments you want to make. But what you should do is to write to the firms I have named below for their price-lists of chemicals.[1] These will include not only all those I have named in this book but many others. Many of the chemicals you will want can be bought of your druggist, and you can often get test tubes and other ordinary pieces of chemical glassware of him.

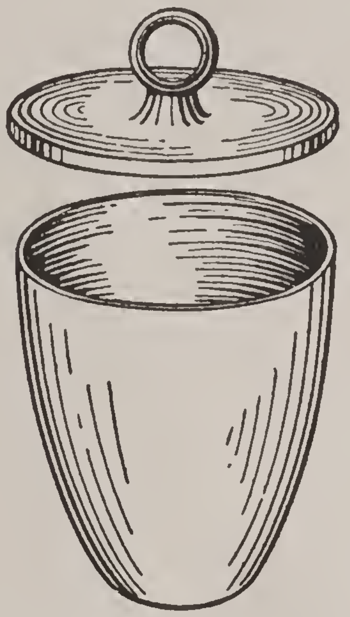

FIG. 23.—A Porcelain Crucible.

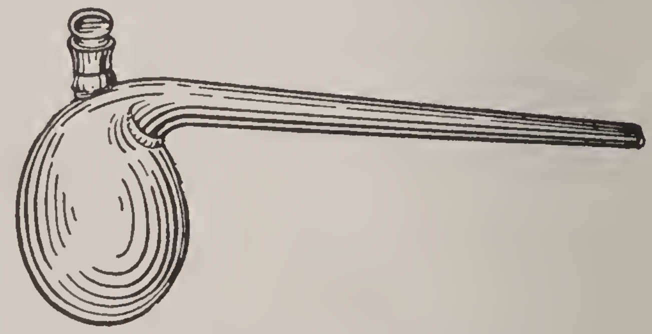

FIG. 24.—A Glass Retort.

While many of the chemicals come in the form of solid substances, it is a good scheme to keep them in tightly corked bottles, while acids and other liquids should be kept in glass-stoppered bottles. Label each bottle carefully and then place all of them in a cabinet where the light cannot reach them, as some compounds decompose under its action.[2] Place them in the cabinet in alpha-

[1] When you have made out the list of the chemicals and chemical apparatus, you should write to the L. E. Knott Apparatus Co., Cambridge, Mass., or to Eimer and Amend, Third Ave. and Eighteenth St., New York City, and either firm will not only quote you prices but will give you other information you want.

[2] See Chapter XII.

betical order so that you will know at a glance just what you have in stock and, what is equally to the point, will be able to find the one you want without having to hunt for it. It is also a good plan to keep all the poisonous chemicals together and to put a red label on each of these, and tie a string around the neck of each one, to the end that you will not mistake them.

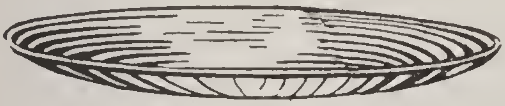

FIG. 25.—A Watch Crystal.

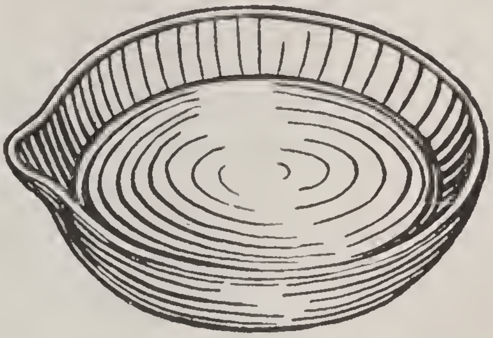

FIG. 26.—A Porcelain Evaporating Dish.

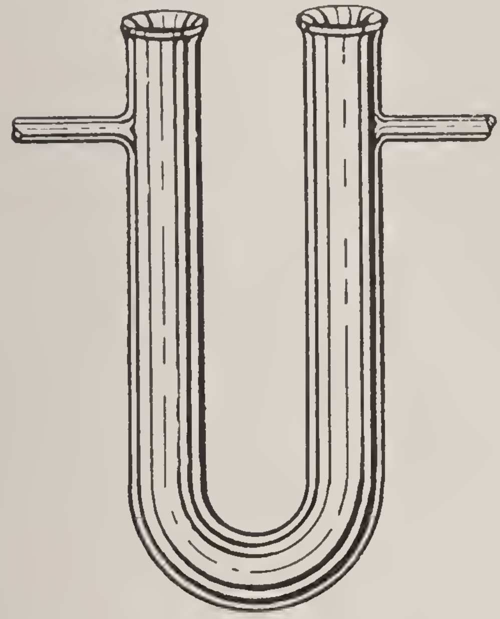

FIG. 27.—A Five-Inch U-Tube.

FIG. 28.—A Pair of Tweezers.

Indicator Papers and Solutions. *Indicators* are papers or liquids that change from one color to another when they are dipped in or mixed with an acid or an alkaline solution. Hence, if you want to know whether a solution is acid, neutral, or alkaline you have only to test it with an indicator.

How Litmus Paper Acts. Litmus is a blue coloring matter that is found in *lichens*, and hence its name. When this is extracted from the plant it is dissolved in hot water (H_2O) and a sheet of absorbent paper, that is, paper without any sizing in it,—like filter paper—is dipped in the solution and then dried, taking on a *blue* color; this, then, is the way that *blue litmus paper is made.* If, now, you dip the paper in a weak acid solution it will turn red, and this gives you *red litmus paper*; finally, if you dip this in an alkaline solution it will turn blue again. Litmus paper is the simplest, though not the most sensitive, of the indicators, but it will serve for all your experiments. You can buy blue and red litmus paper already prepared, as well as the other indicators that follow.

How Phenolphthalein Acts. Phenolphthalein ($C_{20}H_{14}O_4$) —pronounced *fen-ol'-tha-lein*—is a colorless substance that is much used by chemists as an indicator. This is because it is very sensitive to acids, even to those of the weakest kind, but it shows the presence of an alkali only when the latter is strong. It acts the reverse of litmus in that it remains colorless when it is used with acids, but turns red in alkaline solutions.

How Methyl Orange Acts. This indicator is a complex compound with the gentle formula of ($CH_3NC_6N_4 . N. C_6H_4SO_3N$) and you will come to a full realization of what this means as you get farther along in the book. It is a vegetable substance, and when added to an acid solution it turns *red*, and when added to an alkaline solution it turns *yellow*. Its value lies in the fact that it is a very sensitive indicator of weak alkalis.

How Congo Red Acts. Congo red ($NC_{32}H_{22}N_6S_2O_6$) has also, as you will see from its formula, a complex structure. Congo red paper when dipped in an alkaline solution remains *red*, and when dipped in acid solutions turns *blue*, hence it acts in just the reverse way from litmus paper. Its especial usefulness lies in the fact that it shows graduation of color, by which is meant that the depth of the blue it turns depends entirely on the strength of the acid which you are testing.

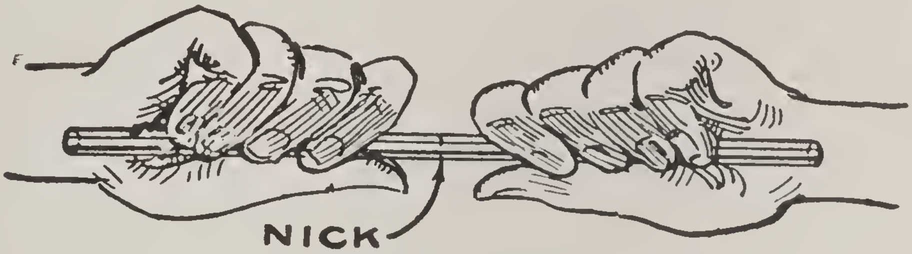

FIG 29.—How to Cut a Glass Tube.

How Sulphide Test Paper Acts. Sulphide test paper is used for testing the presence of sulphur (S) in water (H_2O) and other liquids. If the liquid contains sulphur (S), the test paper will turn brownish-black.

How to Work Glass Tubing. While you can buy glass tubing bent to whatever shape you want it, still in making chemical experiments you will often find it a great convenience to be able to bend it yourself, and as it is an easy and pleasant job, I'll tell you how to do it.

How to Cut a Glass Tube. As I have mentioned under the caption of "The Apparatus You Need," you want to get the kind of glass tubing that goes under the trade name of *German soft glass*, as this kind melts at a comparatively low temperature. This kind of tubing comes in 2-foot lengths, and so the first thing is to know how to cut off a piece of the length you want to use. To do this, you need only to file a nick in it with a three-cornered file, when it will easily break in two. The way you hold it to do this is shown in Fig. 29.

How to Smooth up the Sharp Edges. Hold the end of the tube that you want to smooth up in the flame of your alcohol lamp, or, better because it is hotter, a Bunsen burner, and turn it rapidly around. As soon as it is hot enough it will begin to melt, and this will round the sharp edges, and you can easily see just how far to carry the process.

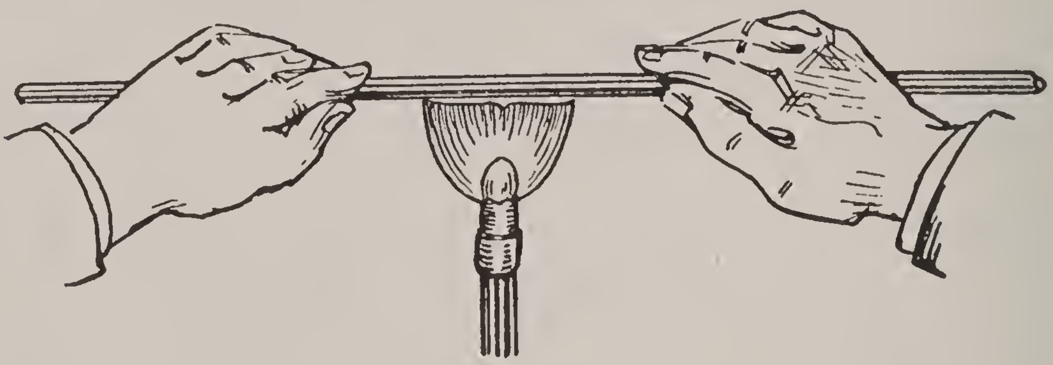

FIG. 30.—How to Bend a Glass Tube.

How to Bend Glass Tubing. To bend a piece of glass tubing, hold it in the flame of your alcohol lamp or Bunsen burner with a wing-top attachment until it is red-hot an inch or so on either side of the place where you are going to make the bend. Now turn the tube rapidly round in the flame so that it will heat equally all over until it gets soft. This done, take it out of the flame and while it is still at a red heat you can bend it to whatever shape you want. You must not heat it until it is too soft, or you will find that the walls of the tube will collapse and so close up the bore. Figure 30 shows the way it is done.

How to Draw a Glass Nozzle. In many cases where you want a tube with a nozzle on it you can take the rubber bulb off a pipette, that is, a medicine dropper, and use it, but if you should want a nozzle with a little larger or a little smaller opening in it, the only way to get it quickly is to make one. To do this, cut off a piece of glass tube about 4 inches long, hold both ends of it so that the middle will be in the flame of your lamp or burner, and at the same time keep turning it round so that it will heat evenly.

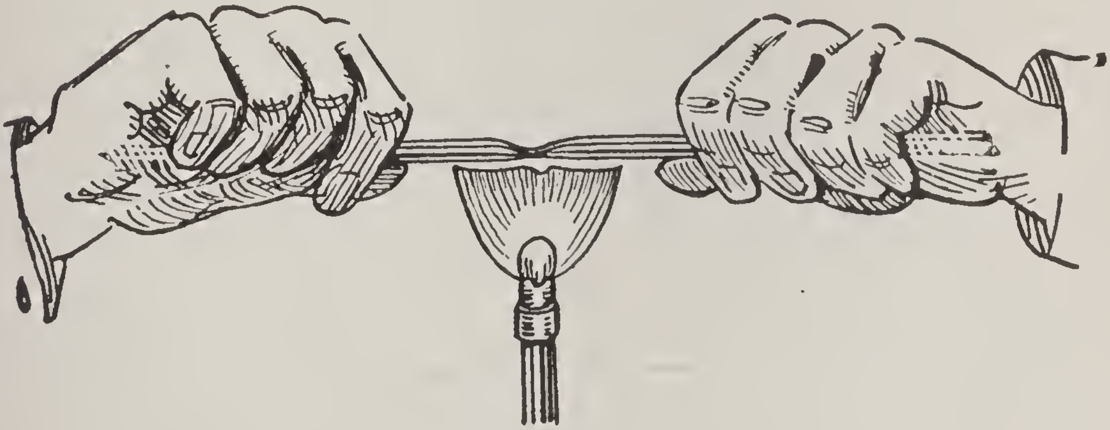

FIG. 31.—How to Draw a Glass Nozzle.

When it gets red hot it will be quite soft, and you can then draw it out until it is very thin in the middle, as shown in Fig. 31. Now take it out of the flame and make a cut with a file at the place where it will give an opening of the size you want; this done, gently tap it with your file. It will break off and the nozzle is ready to use.

CHAPTER II.

AIR, THE MIRACLE-WORKER

The word *atmosphere* is generally used to mean the whole mass of the air that surrounds the earth, while the word *air* itself is taken to mean some small or large part of it. The atmosphere is really a shell formed of gases, the inside of which has, naturally, the same shape as that of the exterior of the earth, while its outside surface is more like a football than it is like a baseball, since it is flattened out on its opposite sides, as shown in the diagram, Fig. 32.

The Height of the Atmosphere. To just what height the atmosphere extends is not known with any degree of certainty, but it is variously estimated to be from 50 to 200 miles, and it may extend in a highly rarefied state even farther than the last-named figure indicates. Now there are several ways by which its height can be calculated, but all of them give results that are different, as the following will show.

Since the speed of the earth as it turns on its axis is known, its *centrifugal force*, that is, the force by which a rapidly rotating body tends to throw things off from its surface, can be figured out. This action is, of course, opposed by the *force of gravity*, which tends to hold all things down to the surface. But as gravity exerts a constantly decreasing pull on the air the higher up it gets, it can be shown that the upper limit of it is only about 50 miles.

Another way to find the approximate height of the upper limit of the atmosphere is with a *barometer*. This is an instrument that really measures the weight, or pressure, of the air, and as this decreases with its height it is easy to calculate its upward limit by taking several readings of the barometer at different levels near the surface of the earth. This method shows that the height of the atmosphere is about 100 miles.

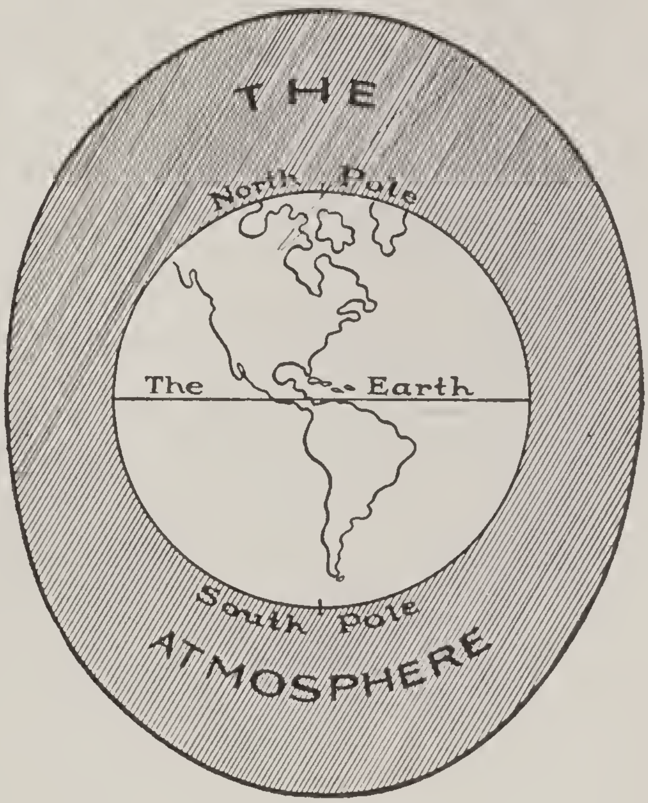

FIG. 32.—The Earth's Atmosphere is Shaped Like a Football.

A third and very interesting way to find the height of the air, is by the *length of twilight*. If the earth were not surrounded with an atmosphere, there would be no twilight and the day would suddenly change into night the moment the sun sank below the horizon. But if the earth's atmosphere reached a height of a thousand instead of a hundred miles or so, then we should have daylight all the time, for the light of the sun is refracted, that is, it is bent out of its course, and diffused, or spread out by the gases of which the former is composed, and so some of the light from it would reach us whatever the relative positions of the earth and sun might be. Hence, it is the height of the atmosphere that determines the length of our twilight.

Thus while the sun is yet below the horizon in the morning, its rays are bent up and we get some of its light when we have dawn, and again when the sun drops below the horizon in the evening, and then we have twilight.

In the extreme northern and southern regions where the days are the longest, twilight is always present, so that there is enough light to see by throughout the whole night. Oppositely, at the equator twilight is very short, and on top of the Andes it lasts for only about half an hour. By figuring the height of the atmosphere on the basis of the length of twilight, the results show that the limit of the atmosphere is reached in the neighborhood of 200 miles above the earth's surface.

The Weight or Pressure of the Air. The atmosphere must have weight or else it would not cling so tightly to the earth's surface, and since it has weight it must exert a pressure on the surface of the earth. It must be clear, too, that having weight the atmosphere is denser directly on the surface of the earth than it is at the upper levels, in fact a cubic foot of it weighs 1¼ ounces at sea level, while its pressure at sea level is, roughly, 15 pounds to the square inch.

Experiment to Show that the Air has Weight. Here is an experiment that not only shows that the atmosphere has weight but also that it has pressure, and that this is equal in all directions. Take a tumbler that has a flat rim and fill it full of water (H_2O). Then lay a piece of cardboard on top of it and then invert it, that is, turn it upside down, as shown in Fig. 33. While the water (H_2O) weighs so much more than a like volume of air, the pull of gravity on the water (H_2O) will not make it run out, because the force is less than that of the pressure of the air on the surface of the cardboard.

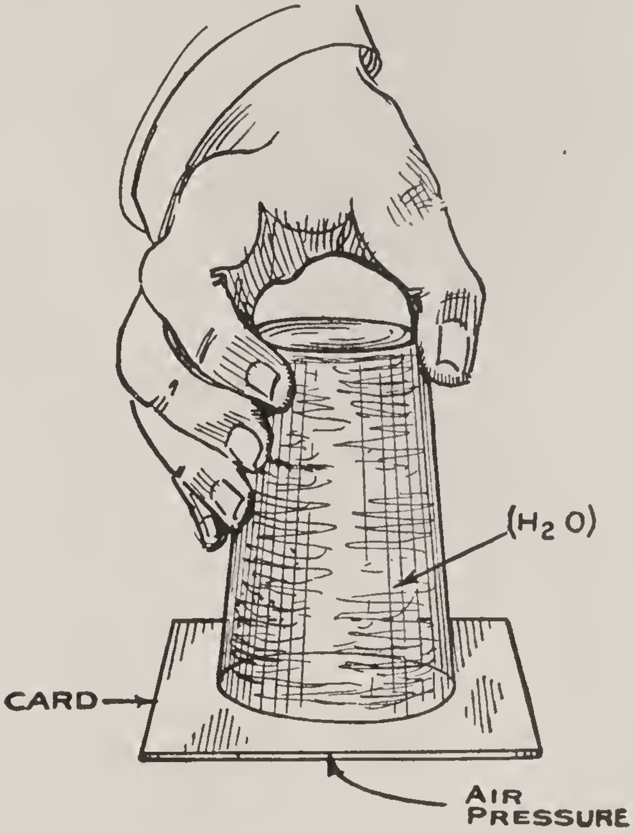

Fig. 33.—An Experiment which shows that the Air has Pressure.

What an Element Is. An *element* in chemistry is a form of matter that cannot be changed into any simpler form.[1] A *substance* is a mass of matter that is made up of one or more elements. Some substances are made up by *merely mixing* two or more elements together *mechanically*, and others are made up by *combining* two or more elements *chemically*.

Experiments to Show what a Mechanical Mixture Is. The two following experiments will give you a very good idea of what a mechanical mixture is and why. Take some *very fine* iron filings (Fe) and an equal amount of *pulverized sulphur* (S) and stir them well together, as though you were going to make a pie. The mass will then take on a greyish color. There has been no chemical action between the two elements, and to prove that they are simply mixed together you have only to pour them out on a sheet of paper

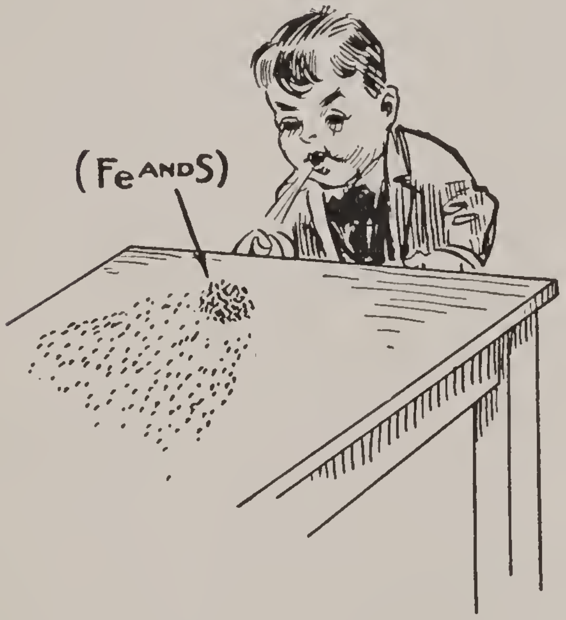

FIG. 34.—Separating Iron Filings from Sulphur by a Stream of Air.

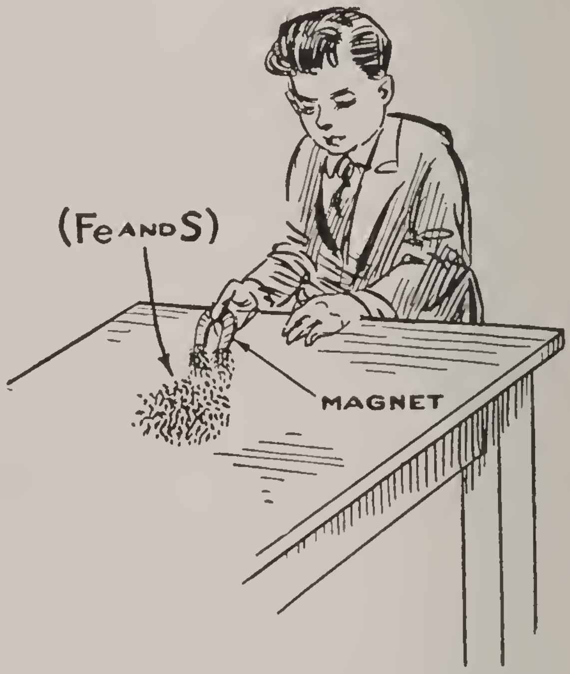

FIG. 35.—Separating Iron Filings from Sulphur with a Magnet.

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[1] The splitting up of the atoms of various elements by Rutherford modifies this statement somewhat, but it still holds good for all ordinary purposes.

and then blow gently on them, as shown in Fig. 34, when the sulphur (*S*), which is very much lighter than the filings, will be carried away, and the iron (*Fe*) only will remain.

A somewhat more scientific test is to hold a magnet close to the mixture, as in Fig. 35, when the iron (*Fe*) filings will be attracted to the poles of it and the sulphur (*S*) will be left behind. The above experiments show clearly enough that the mixture is a purely mechanical one.

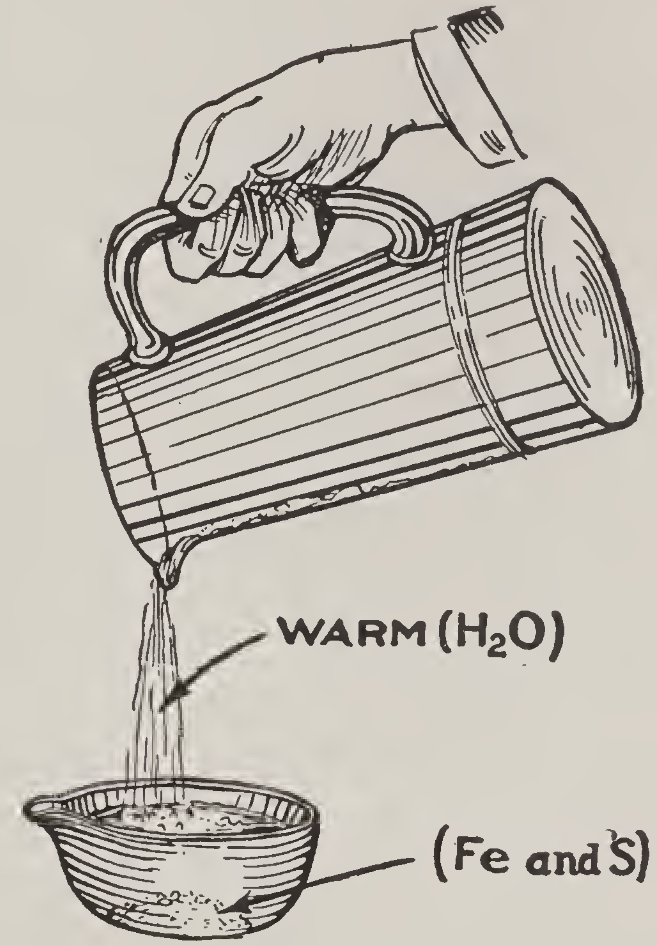

FIG. 36.—Making a Chemical Compound.

Experiments to Show what a Chemical Compound Is. The following experiment will demonstrate in a striking manner what a chemical compound is. Put the iron (*Fe*) filings and the powdered sulphur (S) in an earthenware dish and then pour a little warm water (H_2O) over them, as shown in Fig. 36. In a short time you will see that a change is taking place in the mass, that it gets very hot, swells up, and takes on a black metallic-looking color. The resultant mass has none of the characteristics of iron (Fe) or of sulphur (S), but is a different substance entirely from either one of them, for they have combined chemically and now form ferrous[1] sulphide (FeS).

Now gases behave like solids in that they can either be merely mixed or they can be combined chemically, in which case they will form a new substance. Air is formed chiefly of two gases, which are *oxygen* (O) and *nitrogen* (N), and these are *mechanically mixed* in about the proportion of 23 parts of the former to 77 parts of the latter by bulk, or volume, as it is called, and mixed with these are several rare gases which include argon (A), neon (Ne), krypton (Kr), and xenon (Xe)[2]. These elements of the air are called its fixed constituents, because they are always present in it in exactly the same proportions.

Then the air contains certain other elements and substances, the foremost of which is carbon dioxide (CO_2) or carbonic acid gas, as it used to be called, though incorrectly, for it has no acid properties, water vapor (H_2O) and ammonia (NH_3), and these are known as its variable constituents. Besides these elements and substances of and

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[1] The Latin name for iron is *ferrum*, and from this we get the words ferric and ferrous. The word ferric is used to show that the combining power or valance as it is called is lowest, and ferrous is used to show that it is highest.

[2] Pronounced zeʹ-non.

in the air there are dust, bacteria,[1] and yeast spores floating around in it.

Carbon dioxide (CO_2) is a colorless, odorless compound that is heavier than the air and is formed by the chemical combination of carbon (C) and oxygen (O), and it is this gas that is used to make soda water fizz and sparkle and to cause bread dough to rise. The water vapor (H_2O) consists of oxygen (O) and hydrogen (H) chemically combined, as you will see in Chapter V, while ammonia (NH_3) is nitrogen (N) combined with hydrogen (H).

What the Air Is Good For. Air as a *physical* substance is used at *atmospheric pressure*, that is, just as it is, as a medium for flight by both the winged animals and man, it is also used in the form of *compressed air* and in a *rarefied state* for various industrial as well as experimental purposes. Its first and chief use as a *chemical* substance is in supporting animal life, and the second in supporting combustion, but in both of these cases it is only the oxygen (O) it contains that is used, the nitrogen (N) merely serving to dilute and to spread it. Air can also be liquefied by extracting the heat of it, and liquid air is largely used for experimental purposes.

About Burning and Combustion. When any element or compound combines violently with oxygen (O) it generates heat and often gives off light, and we call this action *burning*, and the process is known as *combustion.* Oxygen (O)

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[1] Bacteria, which is the plural of bacterium, or microbes, as they are popularly called, are vegetable organisms so small that they can only be seen with a high-power microscope. Nearly all of them are harmless, but a few of them are the cause of various diseases.

then, supports combustion, but while it will combine with other elements to make them burn it will not, strange as it may seem, burn itself, and it is well it will not, for otherwise the world would have been consumed in the making.

The reason, then, that you have to supply air in large quantities to make fuel burn is to give the hydrogen (H) and carbon (C) in it plenty of oxygen (O) to combine with, and this is why you blow on or fan a freshly started fire to make it burn, and have a chimney to give the stove, fireplace, or furnace, a draft.

What Rusting, or Oxidation Is. When an element or a substance *unites slowly* with oxygen (O), the action is called *rusting*, or *oxidation*. Here are some simple experiments which show how iron (Fe) and other metals rust.

Experiment to Show how Iron Rusts. Take a piece of iron (Fe) and clean it well, so that there will be no grease on it, or you can file it so that it will expose a clean surface, and then lay it on a damp cloth in the open air. Let it remain over night, and you will then find that the surface of it will be covered with a reddish powder. The iron (Fe) has combined with the oxygen (O) of the air and formed what we ordinarily call rust, but the chemical name of which is *ferric oxide* (Fe_2O_3), and this is sold as *rouge* and *Venetian red*.

Experiment to Show that other Metals Rust. Nearly all the other metals will rust when they are exposed to oxygen (O), but not nearly so quickly as iron (Fe), and so instead of saying that they rust we say that they tarnish. Lead (Pb) and zinc (Zn) will rust in air at ordinary temperatures, but the change is very slow.

Put a small piece of lead (*Pb*) into a porcelain crucible, set it on a stand and place a lighted alcohol lamp, or a Bunsen burner, under it, as shown in Fig. 37; when the lead is melted, stir it with an iron wire and you will soon see a murky yellowish powder appear on top of it; and as you continue to stir in the oxygen (*O*)—for this is what you are really doing—more of the lead is changed into rust, which is chemically called *lead oxide* (*PbO*). Now melt a piece of tin (*Sn*) and stir it in the same way, and a white power, which is *tin oxide* (*SnO*), will be formed.

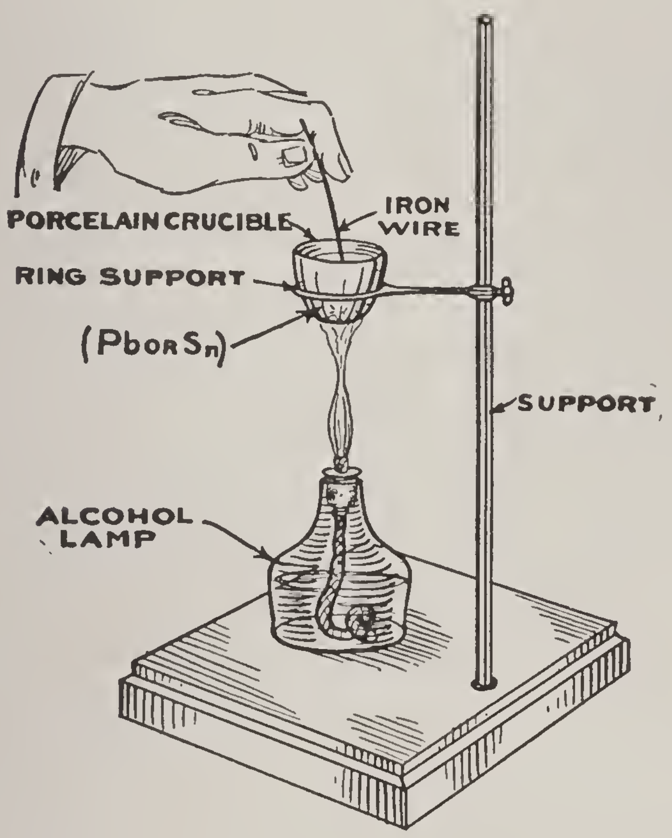

FIG. 37.— Apparatus for Making Lead and Tin Rust.

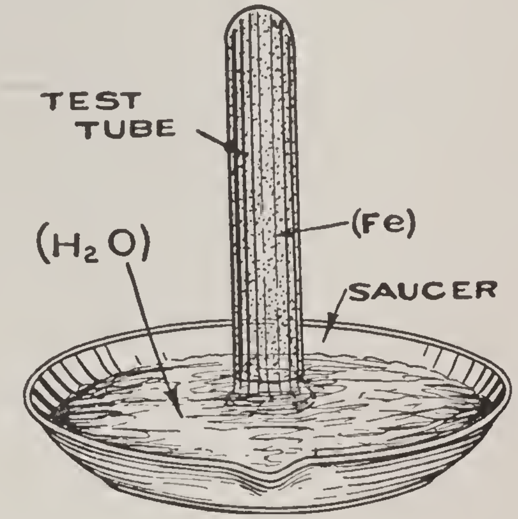

FIG. 38.—Apparatus to Show that Air is Used when Iron Rusts.

Experiment to Show that Air is Used Up when Iron Rusts. This experiment shows that air, or rather the oxygen (*O*) of the air, is used up when iron (*Fe*) rusts.

Take a test tube and dampen the inside surface of it; now put some very fine iron (*Fe*) filings into it and turn the tube over and over until the filings stick to the damp surface; this done, invert the tube, that is, turn it with the open end down, and set it in a saucer of water, as shown in Fig. 38.

The first thing that takes place is that the weight of the atmosphere, or outside air, on the surface of the water (H_2O) in the saucer presses down on it and this forces it up in the tube a little and presses the air that is in the tube into closer contact with the particles of iron (*Fe*). The oxygen (*O*) of the air in the tube makes the damp particles rust and it is thus used up; this leaves more space in the tube so that the weight of the outside air on the water (H_2O) in the saucer presses it still farther up the tube.

In the course of an hour or two, so much of the oxygen (*O*) in the tube will be used up that the water (H_2O) will have reached a height of about ⅕ of the length of the tube. If, now, you will examine the tube, you will see that some of the particles of iron have taken on a brick-colored hue, and this is due to the rust that has formed on the surface of them.

All the particles would rust away if enough air could be supplied to the tube to supply the necessary oxygen (*O*) to them. The reason that only ⅕ of the tube is taken up by the oxygen (*O*) is, obviously, because the other ⅘ is taken up by the nitrogen (*N*), which is not a very active gas.

How Slow Oxidation Causes Decay. Metals are by no means the only elements and substances that are rusted by oxygen (*O*); vegetable and animal matter are likewise affected by it, but in these cases the oxidation is called decay, though the chemical action and the products formed are the same as those produced by burning.

When an animal inhales air, its lungs extract the oxygen (O) from it and deliver it to the blood, the red corpuscles of which carry it to all the tissues of its body. On coming in contact with its food that has been eaten, it oxidizes it and forms carbon dioxide (CO_2) and water (H_2O), with the evolution of considerable heat. The blood returns the carbon dioxide (CO_2) to the lungs, whence it is exhaled. The water (H_2O) is carried off through the kidneys and the pores, while the heat which is set up is used to warm the body.

The opposite process to oxidation is called *reduction*, and just as oxidation is the result of oxygen (O) combining with a substance, so reduction is the result of hydrogen (H) combining with a substance. In other words, when oxygen (O) combines with a substance, or hydrogen (H) is removed from it, it becomes oxidized, and when hydrogen (H) combines with a substance or oxygen (O) is removed from it, it is reduced.

What Spontaneous Combustion Is. It is possible for a slow oxidation to become so accelerated, or quickened, that enough heat will be generated to make a fire, and this kind of an action is called *spontaneous combustion*. The way in which this takes place will be clear when you know that any body that can be oxidized, as for instance iron (Fe), will set free exactly the same amount of heat when it combines with oxygen (O), no matter how slow or how fast the combining action takes place.

When oxygen (*O*) combines with a substance slowly and the latter has plenty of air around it to conduct away the heat as fast as it is developed, there can be no excess of it stored up. But if the substance contains hydrogen (*H*) and carbon (*C*), that is, if it is inflammable, such as oil-soaked rags, and enough air can reach these rags to set up oxidation, but not enough to carry away the heat as fast as it is generated, then the latter will be stored up until the temperature reaches a high-enough point to cause the rags to burst into a blaze.

Substances that Oxygen Will Not Affect. There are some substances that oxygen (*O*) has no effect on and, hence, these cannot burn. While iron (*Fe*) oxidizes rapidly in moist air, it does not oxidize in dry air, and it is therefore largely used for cooking utensils and other purposes.

Gold (*Au*), silver (*Ag*) and platinum (*Pt*) will not oxidize in air, and they are in consequence widely used in the arts, the first two being especially useful for coinage and for jewelry, while the latter, which has a high melting point and is not affected by acids,[1] is useful in certain chemical operations.

How to Make Ozone. As early as 1785 Marum observed that wherever the sparks of an electrical machine appeared a fresh, penetrating odor, something like that of very dilute chlorine (*Cl*) was produced. In 1840 Schönbein discovered that the odor was due to a gas, and this is called ozone (O_3) from a Greek work which means to smell.

Ozone (O_3) is made by adding an extra atom of oxygen (*O*) to a molecule of oxygen (*O*), which is formed of two

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[1] Except aqua regia, see Chapter XII.

atoms of oxygen (O), and this you will find described and pictured in Chapter X. There are several ways by which oxygen (O) can be changed into ozone (O_3), but the easiest way to do it on a small scale is to set up an electric spark, and this you can do with either a Leyden jar or a spark coil. Ozone (O_3) has powerful oxidizing properties and it is therefore a good bleaching agent and disinfectant.

How to Test for Ozone. Put 1 part of pure potassium iodide (KI), 10 parts of starch ($C_6H_{10}O_5$), and 200 parts of water (H_2O) into a beaker and boil them together for 3 or 4 minutes, and they will form a paste. Spread this on a sheet of writing paper evenly, let it dry, and then cut it into strips, and you will have Schönbein's ozone test paper; place these strips together and wrap them up in waxed paper so that the air cannot act on them. If now you will take out one of the strips and moisten it and then put it in air that contains ozone (O_3), even if this is so weak you cannot smell it, the paper will instantly change to a blue color.

CHAPTER III.

EXPERIMENTS WITH OXYGEN, NITROGEN, AND CARBON DIOXIDE

Knowing now something of the nature of air and the chief gases of which it is formed, the next step is to make a small quantity of each one and perform the experiments described in this chapter. Of course you do not really make the gases, but what you do is to separate them, either from the air, which, as you will remember is a mechanical mixture, or else from other substances with which they are chemically combined, and the latter is generally the easier way.

EXPERIMENTS WITH OXYGEN.

A Simple Way to Make Oxygen. With this simple apparatus you can make enough oxygen (O) to do some pretty experiments with, but, naturally, the effects are not so striking as where larger quantities of the gas are used. Put ½ teaspoonful of potassium chlorate ($KClO_3$) and the same amount of manganese dioxide (MnO_2) into your largest test tube and hold it with your test-tube holder over the flame of your alcohol lamp or Bunsen burner, as shown in Fig. 39, and very soon oxygen (O) will be set free. Now slowly sprinkle a little finely powdered charcoal (C), sulphur (S) and other substances, in turn, into the tube by means of a tin trough and you will get some very pretty effects.

A Way to Make More Oxygen. To make a larger quantity of oxygen (*O*) than is possible with the apparatus described above, you will need a small flask with a tightly fitting cork in it, a ring stand, and an alcohol lamp or a Bunsen burner. Now take a piece of glass tubing about a foot long, heat it in the flame of your lamp, and bend it to the shape shown in Fig. 40. This done, make a hole in the cork and push the short end of the tube into it.

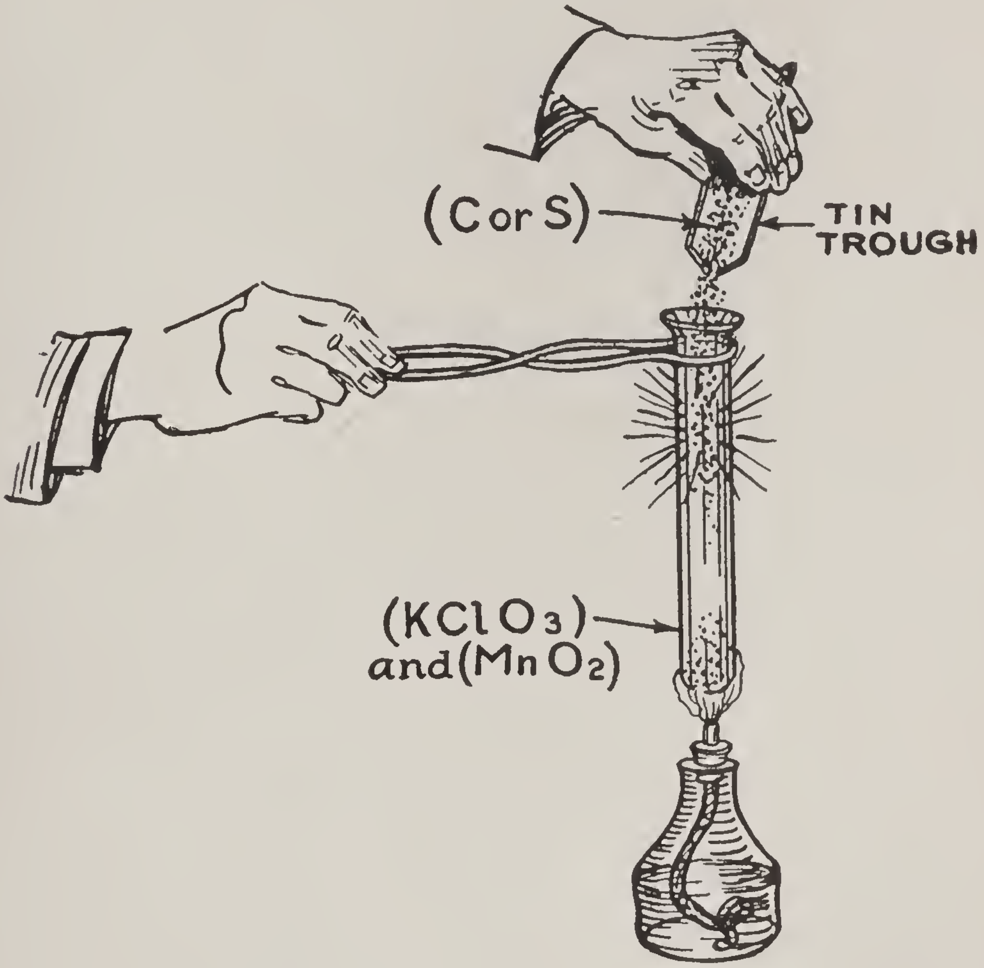

FIG. 39.—How to Make and Experiment with a Little Oxygen.

Next take a large cork and bore a ½-inch hole half-way through it from top to bottom, and bore a ⅜-inch hole through the side of it until it meets the first one, as shown by the broken line in Fig. 41. Push the free end of the tube into this latter hole and set the cork in a glass finger-bowl or other dish. Fill the bowl with water (H_2O) so that it covers the top of the cork, then fill a large test tube with water (H_2O), invert it, and set it on the cork over the hole.

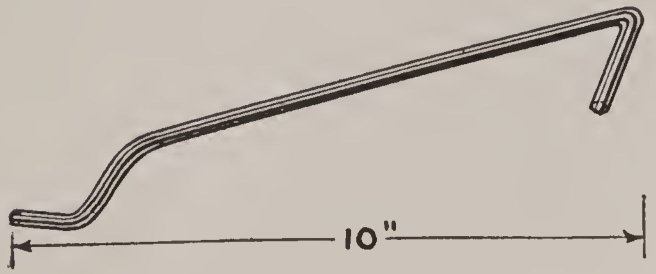

FIG. 40.—How the Delivery Tube is Bent.

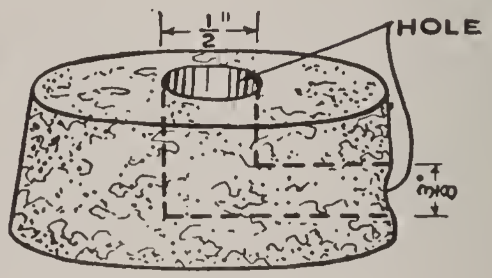

FIG. 41.—The Hole in the Cork.

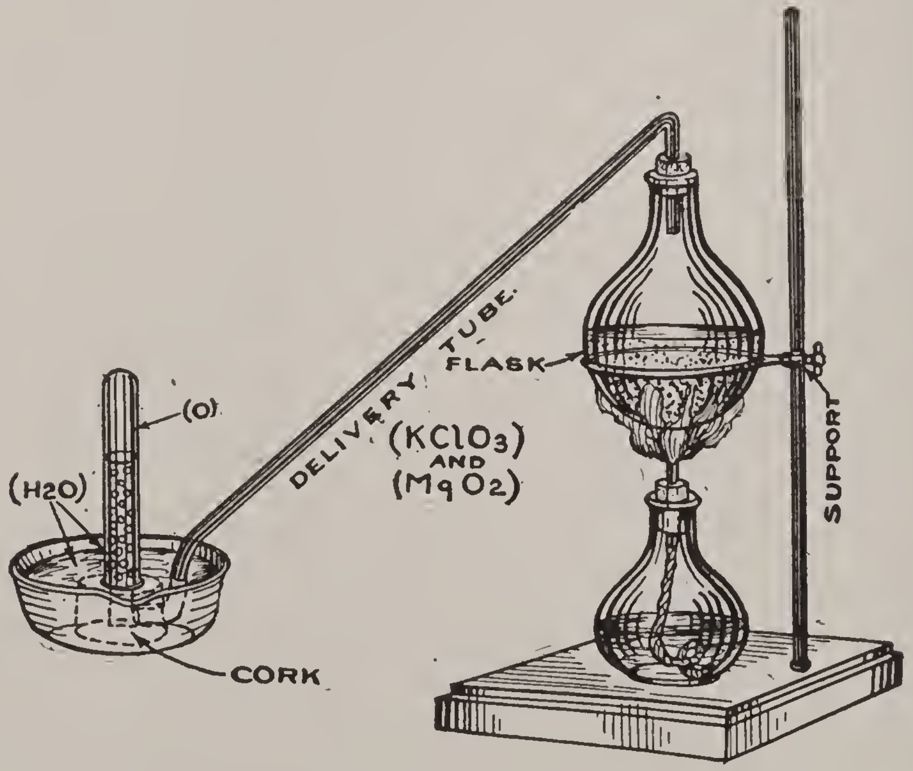

FIG. 42.—The Oxygen Apparatus Complete.

Finally, put 1 ounce of potassium chlorate ($KClO_3$) and 1 ounce of manganese dioxide (MnO_2) into the flask, then fit the cork in tight and set it in the ring of the stand with your lamp or burner under it, as shown in Fig. 42. Light the lamp and the heat will soon act on the manganese dioxide (MnO_2) and set the oxygen (O) of it free, so that it will pass through the delivery tube and up through the water to the top of the test tube. As it gathers there, it will push the water down, and in this way you will know how much oxygen there is in the tube.

How the Experiment Works. When the mixture is heated, the potassium chlorate ($KClO_3$), which contains 39 per cent of oxygen (O), gives up the latter and no action whatever takes place in the manganese dioxide (MnO_2). But if you heat the potassium chlorate ($KClO_3$) without having the manganese dioxide (MnO_2) in contact with it, you will have to bring it to a very much higher temperature before it will be decomposed and liberate its oxygen (O).

Whenever the addition of a substance causes a chemical reaction to take place more rapidly, yet the substance is found at the end of the reaction apparently unchanged, the substance is called a *catalytic agent* and the process is called *catalysis*.

The Self-Lighting Match. Place a large test tube over the delivery tube of your oxygen (O) generator and when it is full of oxygen (O) remove it and hold it with its mouth down. Light a match around which a wire has been twisted, as shown in Fig. 43, and blow it out, leaving only a glowing spark. Now, if you insert the smoldering match by aid of the wire into the tube of oxygen (O), the match will at once ignite again and blaze with more brilliancy than before, as in Fig. 44.

How the Experiment Works. The air, as we know, is oxygen (*O*) diluted with about three times its volume of nitrogen (*N*). The number of particles of oxygen (*O*) in a given volume of air is, therefore, much less than in the same volume of pure oxygen (*O*). When combustion takes place in pure oxygen (*O*), the heat that is liberated is expended in raising the temperature of the oxygen (*O*) alone, and the rapidity of the combustion depends chiefly on the temperature of the oxygen (*O*).

Fig. 43.—
The Wire on the Match.

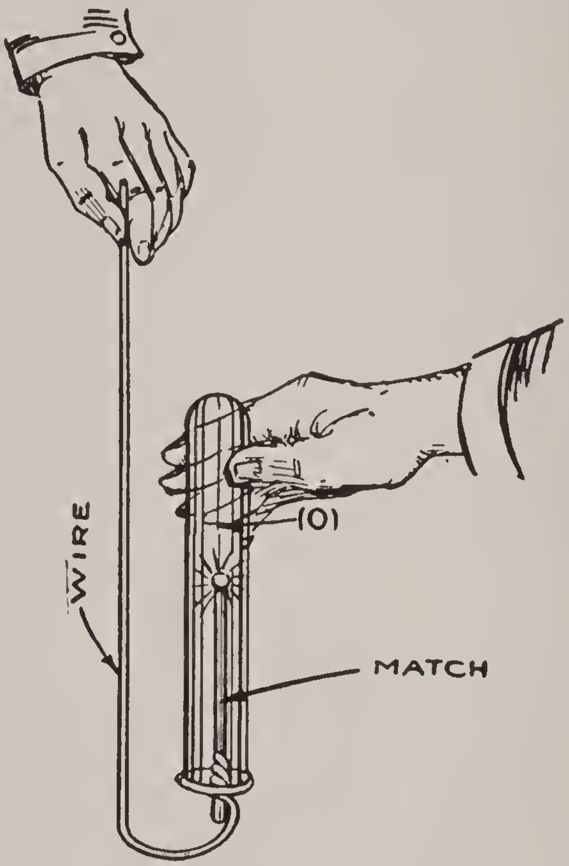

Fig. 44.—
The Self-Lighting Match.

The Flashing Charcoal Pill. For this and the following experiments use a beaker, or a glass tumbler will do, instead of the test tube, so that you can have a larger quantity of oxygen (O) to work with. Take a bit of charcoal, which is practically pure carbon (C), made of bark, or very soft wood, about the size of a pea and fasten the end of a bent wire to it to form a handle.

Soak a bit of cotton in alcohol (CH_3OH) and wrap this around the charcoal pellet. Now light the cotton and hold it in the beaker or tumbler of oxygen (O), and the cotton will quickly burn away and the incandescent charcoal (C) will throw out flashes like an arc light.

How the Experiment Works. When carbon (C) burns in more oxygen (O) than it needs to support combustion, carbon dioxide (CO_2) is formed. You can prove this by moistening a piece of blue lit-litmus paper and, after the charcoal pellet has burned out, pressing this paper against the inside of the beaker, or tumbler. It will turn red. This is because when carbon dioxide (CO_2) is dissolved in water it makes carbonic acid (H_2CO_3), though neither carbon dioxide (CO_2) nor water (H_2O) has any acid property in itself.

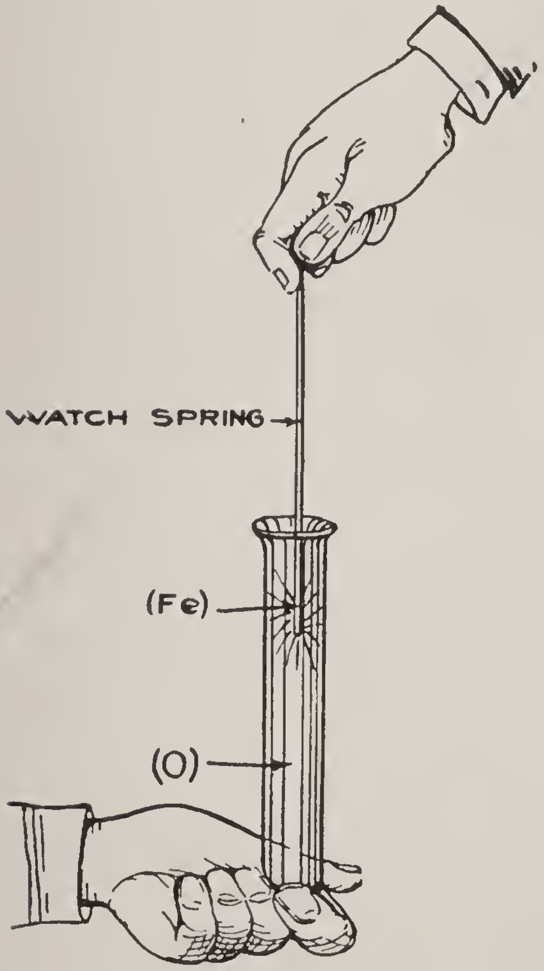

FIG. 45.—The Scintillating Watch-Spring.

The Scintillating Watch-Spring. Take a piece of watch spring (Fe) about 6 inches long and straighten it out by running it between your fingers. Wrap a bit of cotton which you have moistened in alcohol (CH_3OH) around one end and light it. Then hold the steel spring in a test tube or a beaker of oxygen (O) and it will ignite and, once started, will burn with great brilliance, scintillating beautifully, as shown in Fig. 45. At the same time incandescent drops of dross will fall and a reddish vapor will condense on the surface of the test tube or beaker.

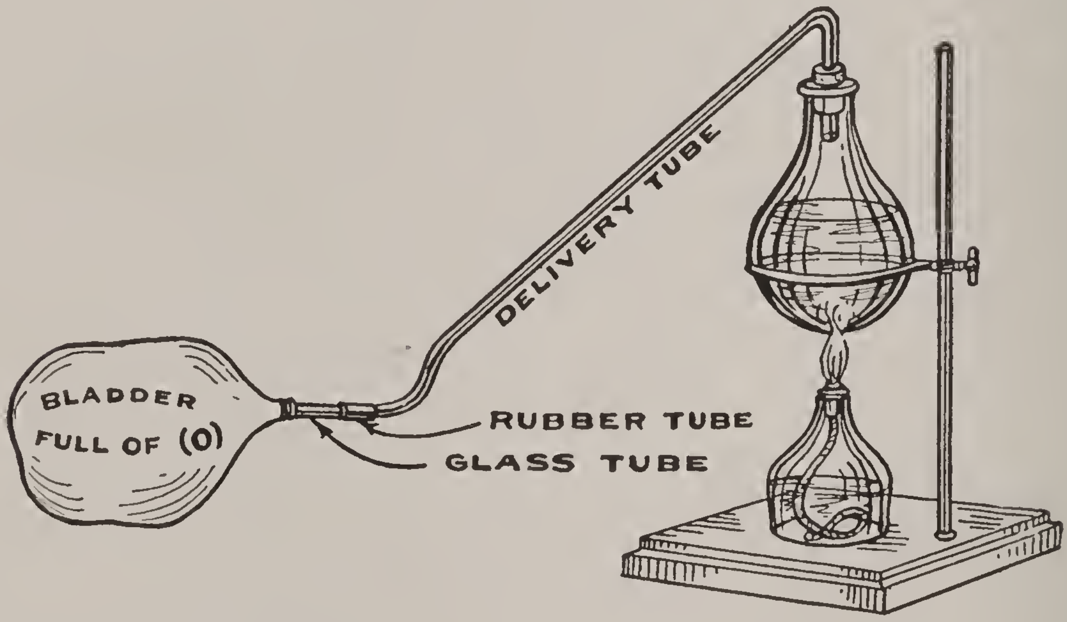

Fig. 46.—Filling the Bladder with Oxygen.

How the Experiment Works. When iron (Fe) burns in oxygen (O) they combine and form ferrous oxide or *oxide of iron* (FeO) as it is called. This substance is a neutral compound, that is, not acid nor yet alkaline, and this you can easily prove with a piece of litmus paper.

The Strange Action of Oxygen on Phosphorus. Connect the free end of the delivery tube of your oxygen-making apparatus, see Fig. 46, with the tube in a toy rubber balloon, or a bladder, then generate enough oxygen (O) to inflate it, as shown in Fig. 46, and close up the mouth of the tube with a bit of wax to keep the gas from escaping. This done, put a few pieces of *red phosphorus* (P), each about the size of a buckshot, in a small beaker and set this in a porcelain bowl, or other deep vessel; now fill both of them full of boiling water (H_2O) and remove the wax from the mouth of the tube and place the tube in the beaker, as shown in Fig. 47. The phosphorus (P) will ignite and burn with exceeding brilliancy under the water (H_2O).

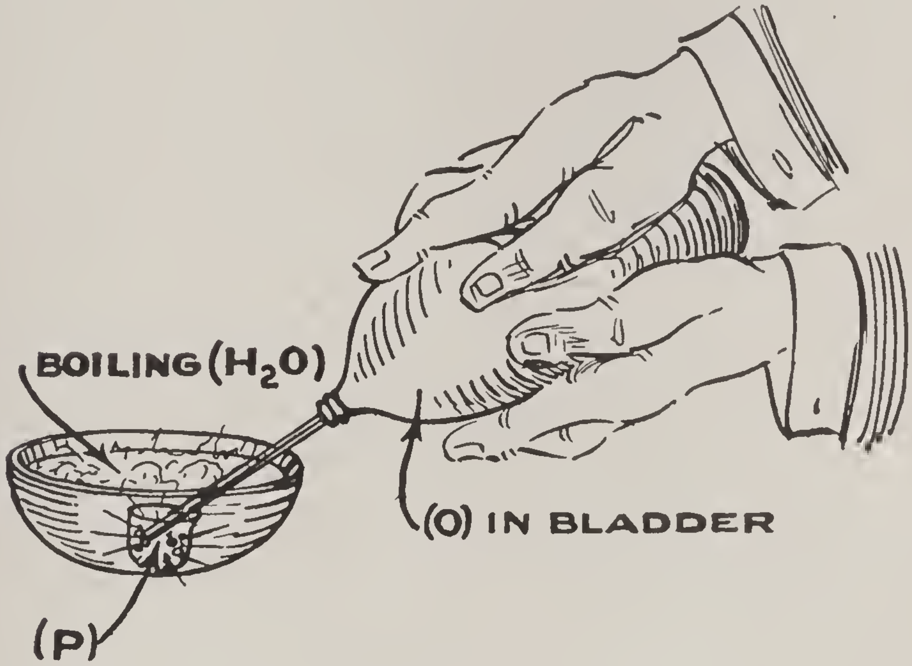

FIG. 47.—Directing a Stream of Oxygen on Phosphorus.

How the Experiment Works. The phosphorus (P) combines with the oxygen (O) that is directed upon it, and the reason it will burn under water (H_2O) is by virtue of the fact that while the hot water (H_2O) melts it, it will not dissolve it, and coupled to this it has an extraordinary affinity for oxygen (O).

NOTE.— In making this experiment, be sure to use red phosphorus instead of *yellow* or *white phosphorus*, as it is not poisonous like the two latter kinds. Also it is a good plan to place a sheet of wire gauze over the bowl while the experiment is in progress.

How to Make an Oxy-Calcium Light. This is a very dazzling light which is also called a *lime-light*, because a piece of lime is used in its production, and a *Drummond light*, after the man who invented it. You can make one on a small scale easily enough, for all you have to do is to drill a 1/16-inch hole in a board and set the end of a stiff wire into it, the height of which is the same as that of your alcohol lamp. Now take a piece of quicklime, that is, calcium oxide (*CaO*), and set it on the free end of the wire support so that it will be in a line with the flame.

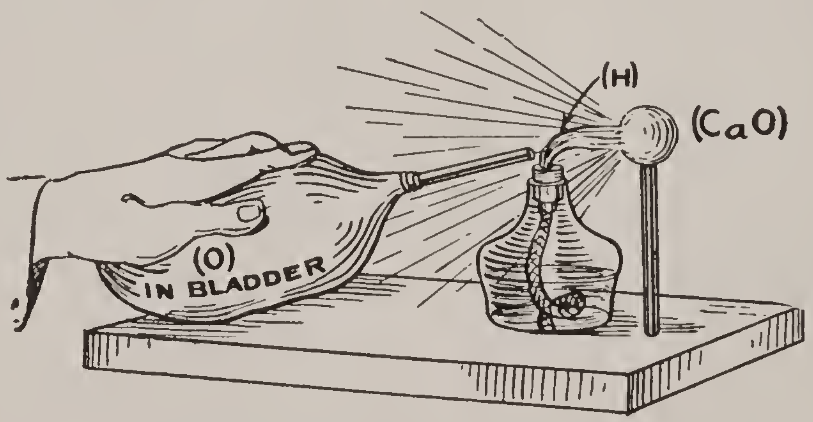

FIG. 48.—How to Make an Oxy-Calcium Light.

The next step is to make enough oxygen (*O*) to fill the toy balloon or bladder with, then press down on the latter and direct the stream of gas on the flame so that they will both strike the piece of lime as shown in Fig. 48. The flame thus produced will be very hot and it will heat the lime to incandescence at the point where it strikes it, and the result is a light of dazzling brightness.

How the Oxy-Calcium Light Works. The chief element in alcohol (CH_3OH) that burns is hydrogen (H). Now when oxygen (O) and hydrogen (H) are mixed together at ordinary temperatures no chemical action takes place; if you seal the mixed gases in a tube and keep this tube heated to 300 degrees for several days, a small amount of the gases will combine to form water (H_2O). At 500 degrees they will combine, though still very slowly; but if you raise the temperature to 700 degrees they will combine instantly and develop an intense heat. Since the alcohol flame is hotter than 700 degrees, the oxygen (O) and the hydrogen (H) combine easily.

How Sulphur Burns in Oxygen. Take a bit of sulphur (S) the size of a pea, wrap one end of an iron wire around it, light it, and hold it in a beaker of oxygen (O). The sulphur (S) burns with a wonderful scintillating flame that is violet-colored; the result of the combustion is that sulphurous acid (H_2SO_3), nitrogen (N) and potassium sulphate (K_2SO_4) are set free and the beaker is filled with the fumes of these substances.

EXPERIMENTS WITH NITROGEN.

There are several ways to obtain nitrogen (N), and among these are to burn phosphorus (P) in air, to pass air over finely divided copper (Cu), and by the evaporation of liquid air. In the first two processes the oxygen (O) of the air is taken up by the phosphorus (P) and the copper (Cu), and this leaves the nitrogen (N) behind. For the production of large amounts of nitrogen (N), liquid air is allowed to evaporate. The oxygen (O) passes off first because it is the lighter gas and the nitrogen (N) is left behind.

A Simple Way to Make Nitrogen. Fill a soup-plate half full of water (H_2O), then light a piece of paper and place it in a beaker, or a tumbler, which you invert and set in the water (H_2O), as shown in Fig. 49. You will soon see that the flame of the burning paper grows more and more feeble, and that the water (H_2O) rises higher and higher in the beaker or tumbler, until it is considerably above the level of that in the soup plate. When the paper will no longer burn, you will have considerable nitrogen (N) in the glass.

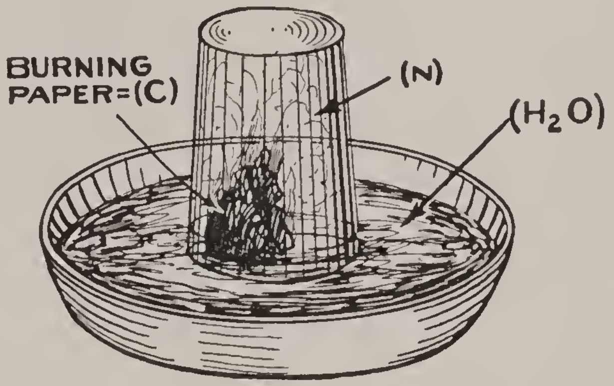

Fig. 49.—A Simple way to Make Nitrogen.

How the Experiment Works. When the paper burns, it consumes the oxygen (O) in the beaker or tumbler, and this produces a partial vacuum in it. The pressure of the outside air on the water (H_2O) in the soup-plate is now greater than that of the air in the beaker, or tumbler, and consequently it forces the water (H_2O) up and into the latter.

The oxygen (O) in the beaker, or tumbler, combines with the carbon (C) in the paper and forms carbon dioxide (CO_2); this gas is heavier than the air and so falls on the surface of the water, which absorbs it; hence the gas that remains in the tumbler is nitrogen (N).

Another Easy Way to Make Nitrogen. Pour a tablespoonful of alcohol (CH_3OH) into a little tin pill-box, or an iron or porcelain dish, having a diameter of 1½ or 2 inches, and set this in a wine-glass. This done, stand the wine-glass in a soup-plate filled with water (H_2O). Now light the alcohol (CH_3OH), and then set a beaker, or a tumbler or a small fruit-jar over them, as shown in Fig. 50.

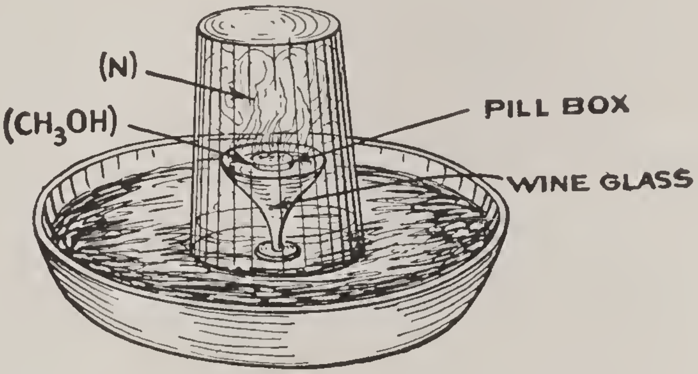

FIG. 50.—A Better Way to Make a Little Nitrogen.

How the Experiment Works. As soon as the alcohol (CH_3OH) is ignited, the oxygen (O) of the air in the beaker, or other vessel, combines with the carbon (C) of it just as it did with the carbon (C) of the paper in the foregoing experiment and forms carbon dioxide (CO_2). The phenomenon of the water (H_2O) rising in the beaker, or other vessel, is due to the same causes as described in the experiment above.

How to Make a Larger Amount of Nitrogen. For experimental purposes you will need considerably more nitrogen (N) than either of the preceding processes will give you. To make a larger amount, place a very little dry *red* phosphorus (*P*) in a small porcelain dish and set it on top of a wine glass; now stand this in a soup-plate filled with water (H_2O) and ignite the phosphorus (*P*) with the end of a wire which you have heated a bright red in the flame of your lamp or burner. Having done this, set a glass jar—a pint fruit-jar will do—over the burning phosphorus (*P*), as shown in Fig. 51. When the phosphorus (*P*) has burned out, the jar will then contain only nitrogen (*N*).

How the Experiment Works. As soon as the phosphorus (*P*) begins to burn, the air in the jar expands because of the heat that it develops, but very soon the air contracts, for the oxygen (*O*) unites with the phosphorus (*P*) and forms a white smoke which is phosphorus trioxide (P_4O_6); this compound falls to the surface of the water (H_2O), combines with it and so disappears, thus leaving only the nitrogen (*N*), which is in a tolerably pure state.

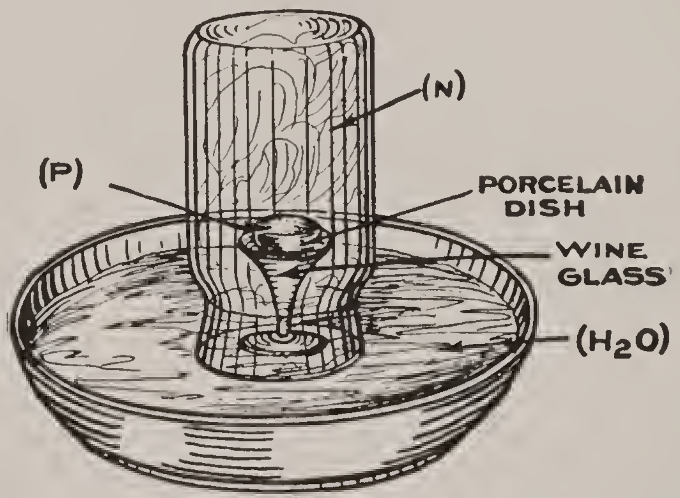

Fig. 51.—How to Make Nitrogen for Experimental Purposes.

The Self-Extinguishing Match. This is the complementary experiment to the one explained under the caption of "A Self-Lighting Match." Twist a copper wire around a match, as shown in Fig. 43, lift the jar of nitrogen (*N*) up from the soup-plate, then light the match and let it get to burning well. When this is done, put it up into the jar, and you will see how quickly the flame will be extinguished. This experiment shows that nitrogen (N) will not support combustion.

What Else the Experiment Shows. The fact that the match will not burn in nitrogen (N) does not at all show that it is a poison. While nitrogen (N) will not support combustion, we take into our lungs a little more than three times as much of it as we do oxygen (O) and with no harmful effect. But carbon dioxide (CO_2), cyanogen (C_2N_2), and several other gases will kill, not only because they cannot support combustion, but because they are poisonous.

EXPERIMENTS WITH CARBON DIOXIDE.

Carbon dioxide (CO_2) is a wonder gas, and many strange and striking effects can be produced with it which are worthy of a place on a magician's program. In this respect it is quite unlike nitrogen (N), which is sluggish and, in consequence, permits but very few experiments to be performed directly with it. Carbon dioxide (CO_2) is a colorless and odorless gas and it is considerably heavier than the air, and for this reason when you perform experiments with it you can handle it just as though it were a liquid like water (H_2O), that is, the vessels need not be corked up, nor covered over, nor inverted, to keep it in them.

How to Show there is Carbon Dioxide in the Air. Fill a clean saucer or a small flat, porcelain dish with clear *limewater*, which is calcium hydroxide ($Ca(OH)_2$), and this you can make by pouring some water on *quicklime*, which is calcium oxide (CaO). Then set the dish of lime-water in the open air, and it will soon be covered with a film of calcium carbonate ($CaCO_3$) the common name of which is *carbonate of lime.* If you will now break the film, it will fall to the bottom of the dish and the operation can be repeated until all the quicklime (CaO) in the solution is changed into calcium carbonate ($CaCO_3$).

How the Experiment Works. That carbon dioxide (CO_2) is present in the air is evident, since this gas must combine with the calcium hydroxide ($Ca(OH)_2$) to form calcium carbonate ($CaCO_3$). Chalk, limestone, marble, egg-shells, oyster shells, coral and pearls, and calcite and Iceland spar are all formed of calcium carbonate ($CaCO_3$).

To Show that You Inhale Oxygen and Exhale Carbon Dioxide. You can easily show that after you inhale oxygen (O) you exhale carbon dioxide (CO_2) by means of this very simple experiment. Fill a tumbler about three-fourths full of lime-water, that is, calcium hydroxide ($Ca(OH)_2$); then take a good deep breath (oxygen (O)) and blow through a straw into the lime-water, as shown in Fig. 52. The clean lime-water will be made murky by the formation of calcium carbonate ($CaCO_3$) in it, and for the same reason as explained under the foregoing caption of "How the Experiment Works."

A better way to make the experiment is to take a wide-mouth bottle and fit it with a cork and two bent glass tubes, as shown in Fig. 53. Pour in enough lime-water, that is, calcium hydroxide ($Ca(OH)_2$), to half fill the bottle, then put the short tube in your mouth and draw in. The outside air will then pass through the long tube and up through the lime-water ($Ca(OH)_2$), which will remain clear. Now repeat the operation, but this time blow through the long tube and the lime-water ($Ca(OH)_2$) will get murky,

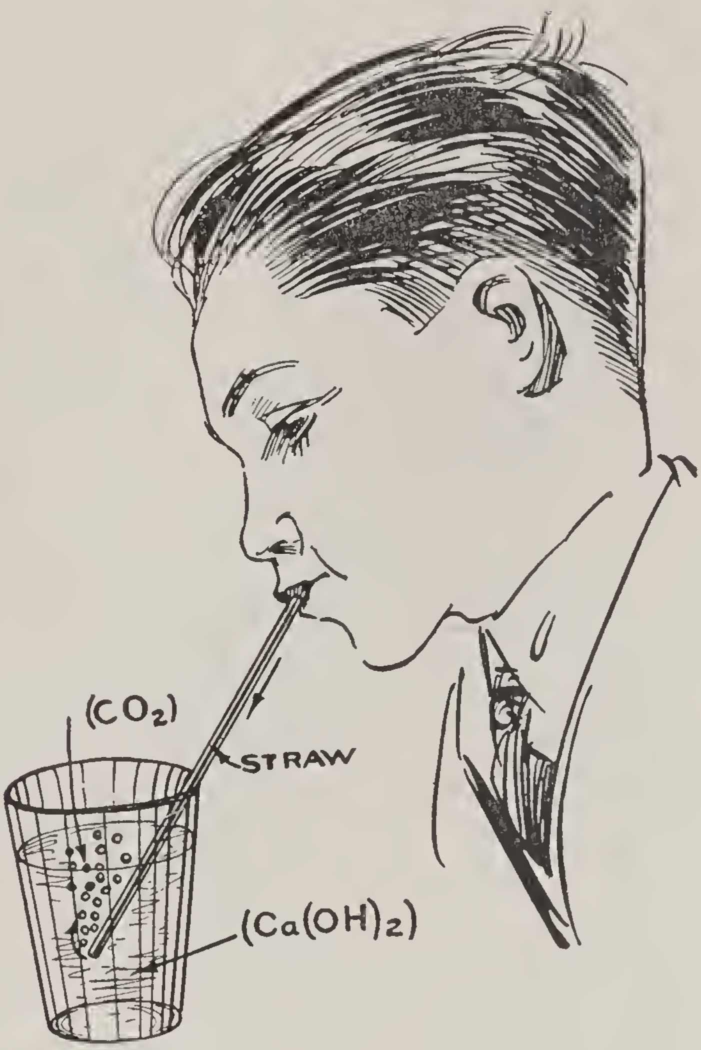

FIG. 52.—A Simple Way to Show Carbon Dioxide.

just as it did in the previous experiment, and for the same reason.

How to Make Carbon Dioxide. In the experiments above, it has been shown that chalk, limestone, and marble are all forms of calcium carbonate ($CaCO_3$) and hence these substances contain carbon dioxide (CO_2). Now all you have to do to make a little of this gas is to put some powdered chalk, limestone, or marble into a test tube half full of water (H_2O), and then add a few drops of hydrochloric acid (HCl) to it, as shown in Fig. 54. Instantly there will be a commotion of the liquid set up by the production of numerous small bubbles of gas which rise to the surface and then break, and these are formed of carbon dioxide (CO_2).

FIG. 53.—A Better Way to Show Carbon Dioxide.

How the Experiment Works. While carbon dioxide (CO_2) is given off in the above reaction, calcium chloride ($CaCl_2$) and water (H_2O) are left behind.

NOTE.— Do not use blackboard crayon for the experiment, as this is usually made of gypsum, that is calcium sulphate ($CaSO_4.2\,H_2O$) and not of chalk, which is calcium carbonate ($CaCO_3$). Since carbon dioxide (CO_2) is a gas that is heavier than air, it will stay in the tube or other vessel when the latter is right side up, just as though you had water (H_2O) in it.

A Better Way to Make Carbon Dioxide. Where you want to make a small amount of this gas for experimental purposes, take a 4-ounce or 8-ounce, wide-mouth bottle with a tight-fitting cork; now bore a hole in the latter and push a piece of glass tube with a ½-inch or ⅝-inch bore—or you can use a test tube from which you have cut the closed end—through it; into the mouth of this fit another cork and push the short end of a bent glass tube with a ⅜-inch bore through it, all of which is shown in Fig. 55. To make the joints gas-tight, put some melted paraffin on them.

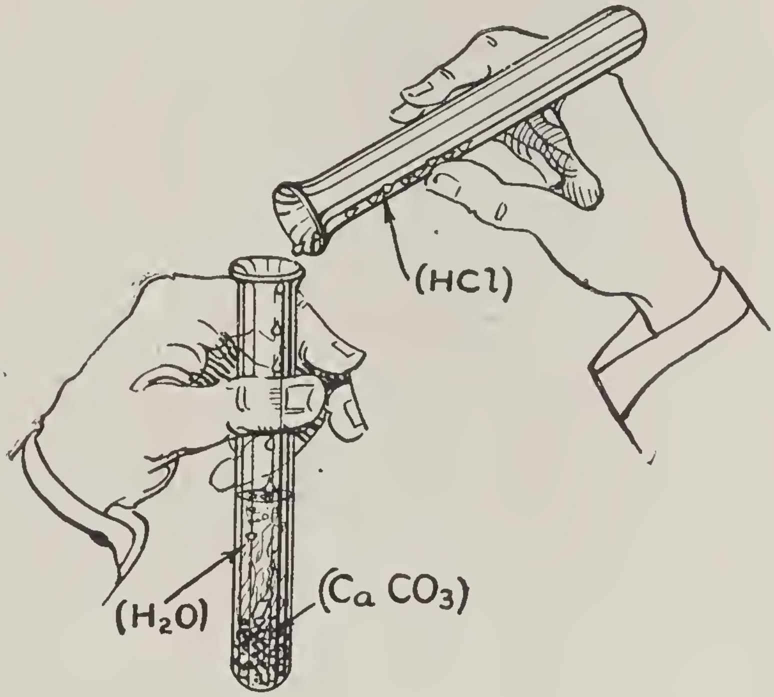

FIG. 54.—A Simple Way to Make Carbon Dioxide.

Use in suitable quantities the same materials called for in connection with Fig. 54.

Where you want a still larger amount of the gas, use a pint jar and fit the cork with a glass funnel tube that is long enough to reach from the outside to the bottom of the jar, and a delivery tube bent as shown in Fig. 56. If you will use powdered marble ($CaCO_3$) instead of chalk you will get a supply of nearly pure carbon dioxide (CO_2). This gas is heavier than air, and has, therefore, a tendency to stay in the bottom of the jar; but as it is set free from the

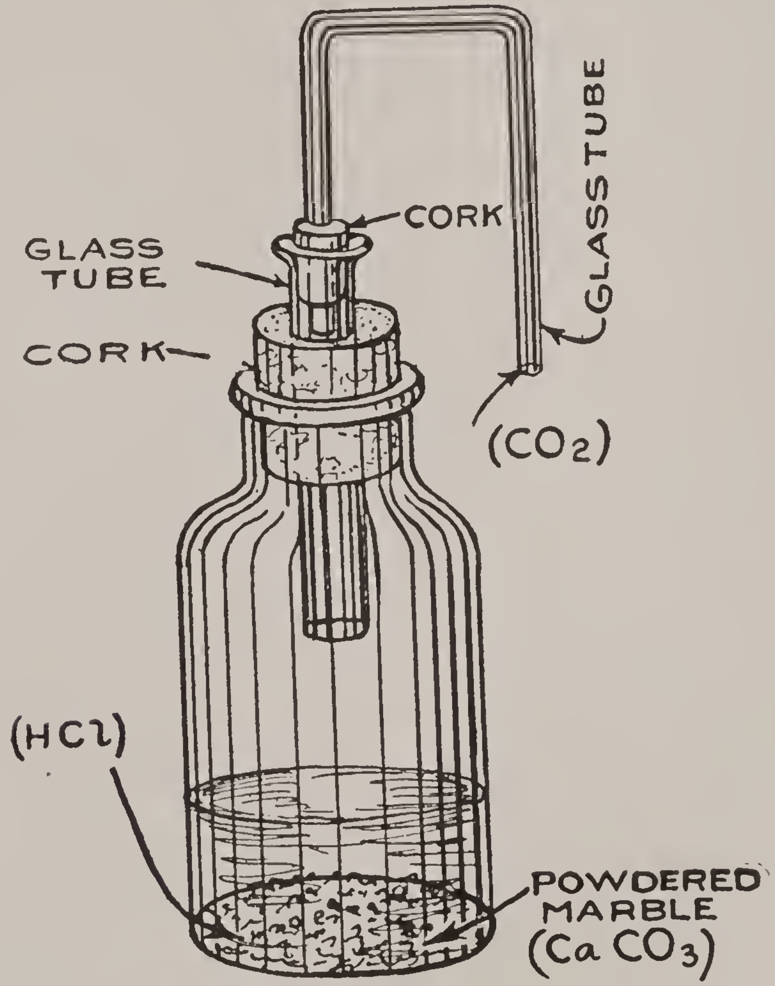

FIG. 55.—A Better Way to Make Carbon Dioxide.

marble in large quantities it is soon under pressure and this drives it out of the delivery tube.

To Show that Carbon Dioxide Will Not Support Combustion. Fill a wide-mouth bottle. or a glass jar, with carbon dioxide (CO_2) and lower a lighted candle into it, and the flame will be extinguished. You can do this experiment as a trick, for to the average person the bottle, or jar, is, to all intents, an empty one; now lower half a dozen lighted candles into the jar one after the other, and the

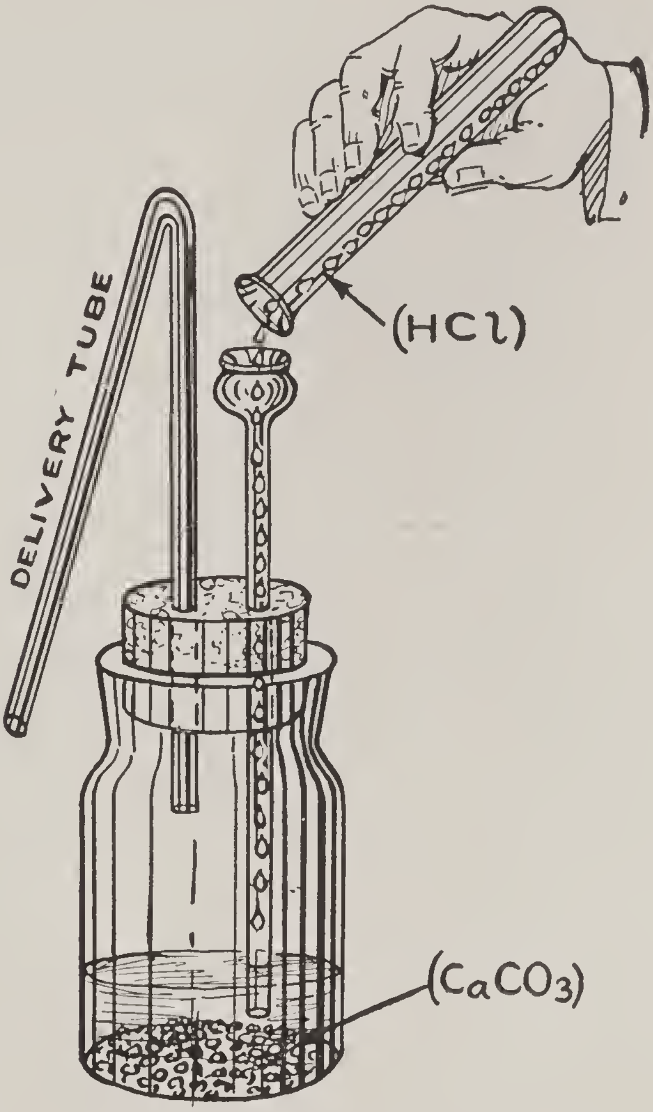

FIG. 56.—To Make a Larger Amount of Carbon Dioxide.

flame of each one will go out as it reaches the surface of the gas. The effect is most mysterious.

Moreover, the smoke from the candle when it goes out does not rise into the air as it is expected to do in the natural order of things but, instead, it floats on top of the unseen gas in a strange and uncanny way, very like a London fog, and if you shake the jar it will set up miniature waves in imitation of the old ocean itself. The reason the smoke clings to the surface of the gas is because it easily mixed with the latter and this holds it down.

To Show that Carbon Dioxide Destroys Life. Carbon dioxide (CO_2) is different from nitrogen (N) in that it kills not only because it cannot support combustion, and, hence, life, but by virtue of the fact that it is poisonous. And yet as high as 6 per cent of it can be breathed without harm when it is mixed with the oxygen (O) and nitrogen (N) of the air. If you are a naturalist as well as a chemist you can kill insects for your specimens and preserve them in their original form and brilliancy of color by simply putting them into a jar of carbon dioxide (CO_2).

A Magical Experiment with Air, Carbon Dioxide, and Oxygen. Take three pint glass jars and let the first one contain ordinary air, fill the second one with carbon dioxide (CO_2), and the third one with oxygen (O), and invert the latter one until you are ready to do the trick. Now wrap a wire around a piece of candle, light it, and then lower it into the jar of air first. Of course, the flame will continue to give its light. Now lower it into the jar of carbon dioxide (CO_2), and the flame will mysteriously go out; draw it out of the jar before the wick cools off and dip it into the jar of oxygen (*O*), and it instantly relights and burns with a dazzling light. Since all the jars are evidently quite *empty*, the average spectator will be at a loss to account for the different actions that take place. The effects are shown in Fig. 57.

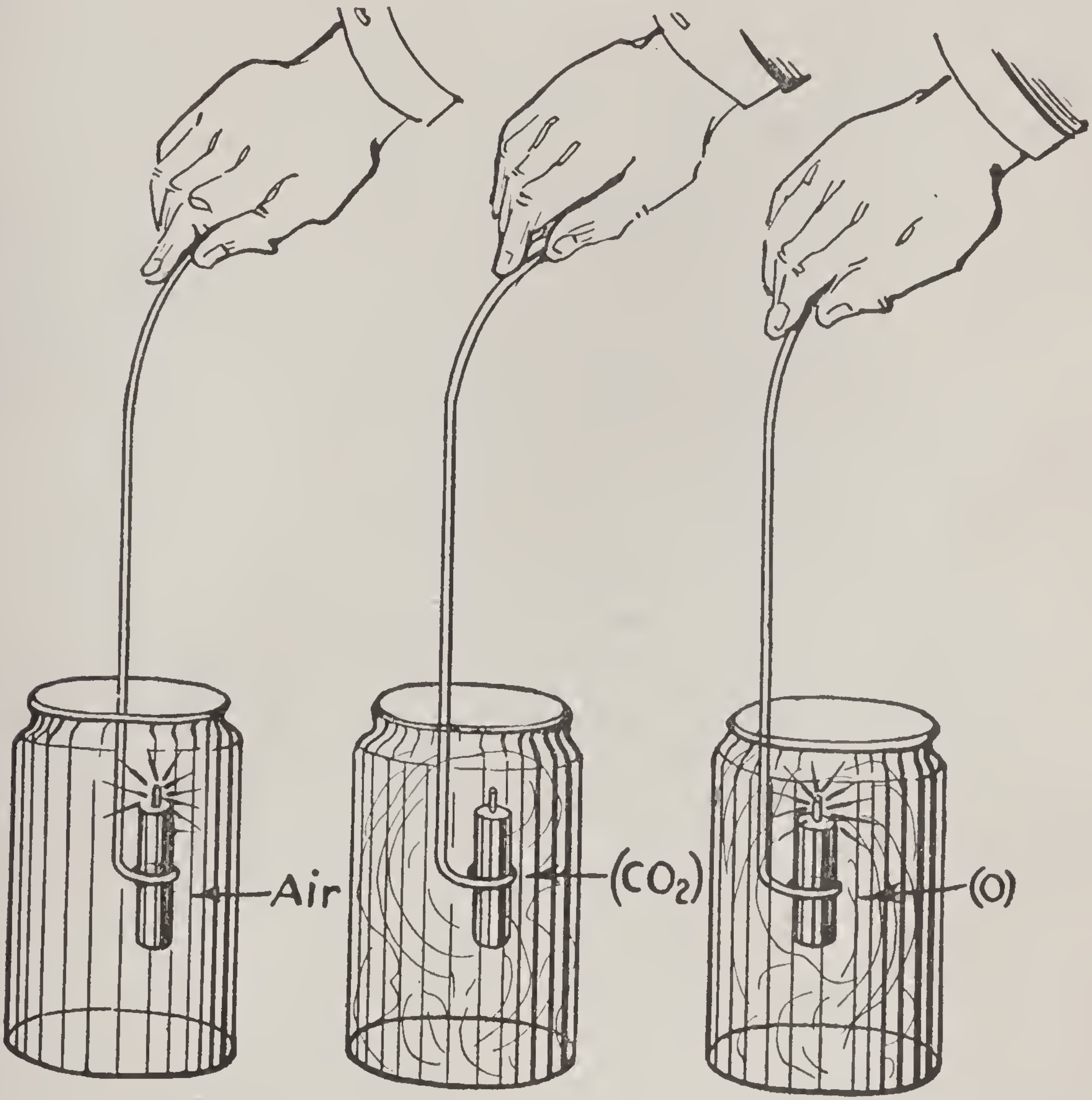

FIG. 57.—A Magical Experiment.

To Show that Carbon Dioxide Has Weight. This is a good magical experiment, too, and for it you need a couple of pint glass jars, one of which you have secretly filled with carbon dioxide (CO_2). Now set a piece of lighted candle in the bottom of the other jar and then pour the contents of the first jar, which is invisible to the spectators but as real as if it were water, into the second jar, as shown in Fig. 58. The candle will be mysteriously extinguished.

To Separate a Candle from Its Flame. Light a candle and lower it into a jar of carbon dioxide (CO_2) far enough so that the tip of the wick is about ½ inch *below* the surface of the gas. The flame, strange as it may seem, will keep on burning above the surface of the gas although it is entirely cut off from the wick.

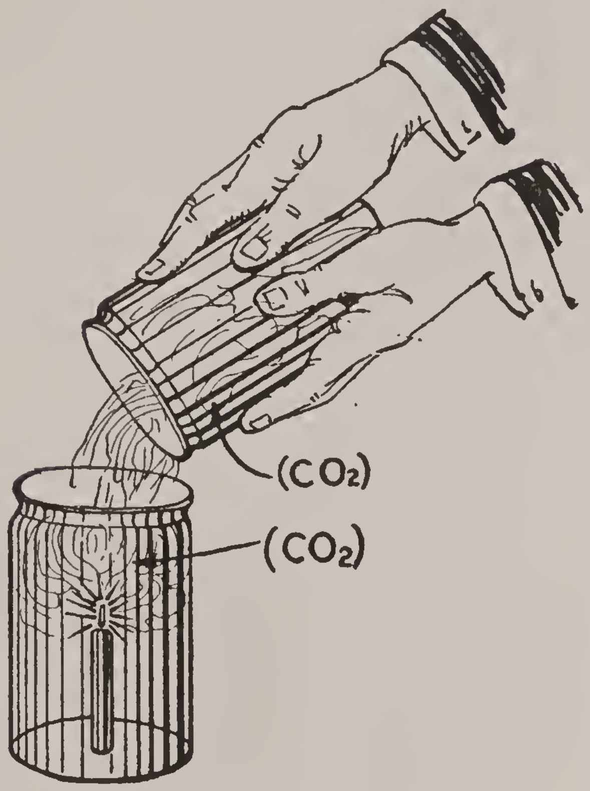

FIG. 58.—Pouring Carbon Dioxide from One Jar into Another.

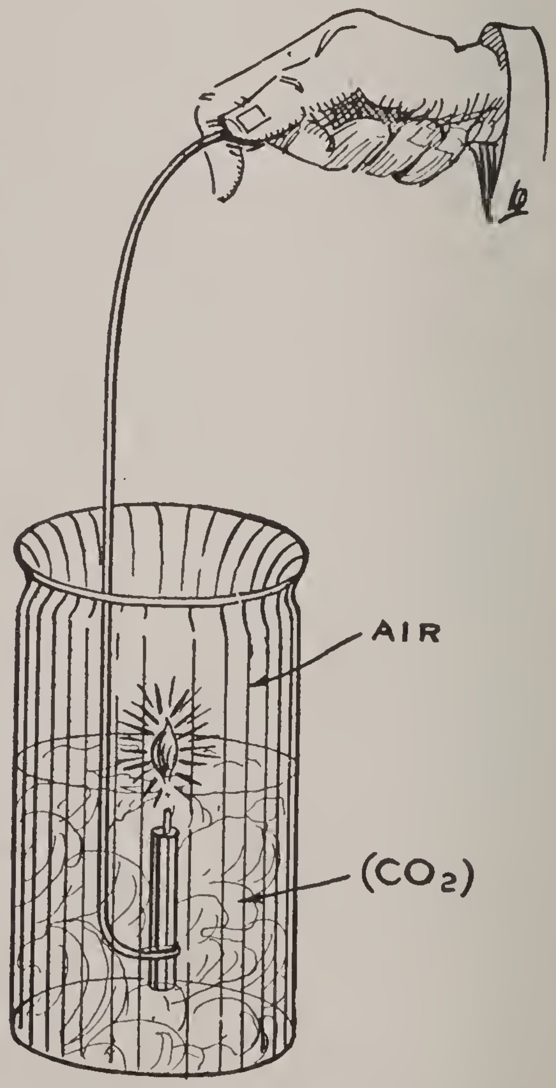

FIG. 59.—Separating the Flame of a Candle from its Wick.

How the Experiment Works. This strange effect, which is shown in Fig. 59, is due to the fact that the heat of the wick lasts long enough to vaporize the paraffin of which the candle is made for a few moments after it is submerged in the gas, the hot vapor from it ascends through the latter, where it is supplied with oxygen (*O*) from the air.

The Levitation of a Soap Bubble. Here is an experiment that would do credit to and gain renown for a Hindu magician. Set a large meat-platter or a tray on the table and fill it to overflowing with carbon dioxide (CO_2). Now blow a large soap bubble with a clay pipe in the ordinary way, hold it over the platter and let it drop on the surface of the gas that fills it. On striking the layer of carbon dioxide (CO_2) it will bounce up and down on the latter like a rubber ball on the sidewalk, and when it finally does come to rest it looks to the spectators as though it were suspended above the platter, or tray, which, of course, it is by the layer of carbon dioxide (CO_2). The experiment is shown in Fig. 60.

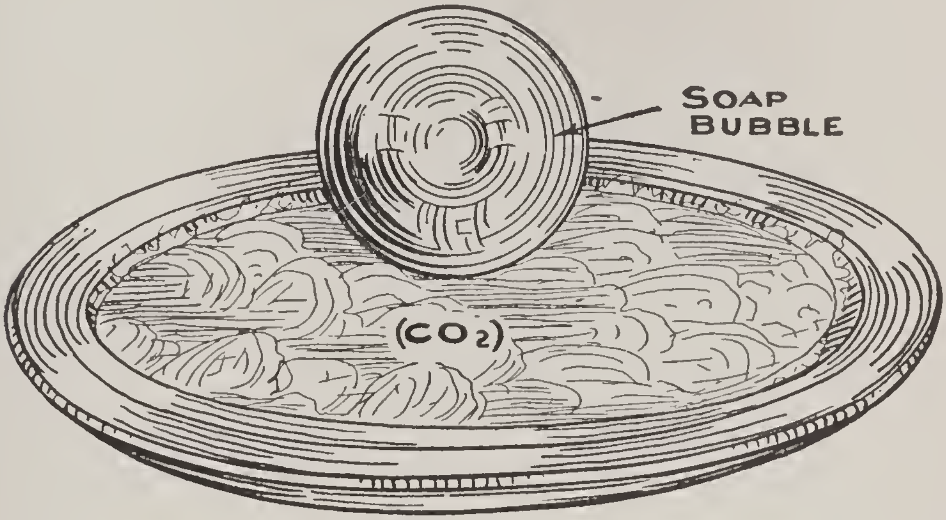

FIG. 60.—The Levitation of a Soap Bubble.

CHAPTER IV.

THE WIZARDRY OF WATER

THE liquid which we call water (H_2O) is as necessary to the existence of living things as air is. Like air, water (H_2O) is formed of two gases, but, differing from air, these are chemically combined and form a liquid nearly 800 times heavier than the former. Water (H_2O) covers three-fourths of the earth's surface, the oceans taking up the larger part of it, and soundings have been made which show that at various points it is more than 5 miles deep. As you have seen in Chapter IV, the air has a large amount of water (H_2O) in it in the form of vapor, and the so-called dry land is saturated with it, while both plants and animals are made up of from 50 to 75 per cent of it, hence without it life could not exist.

Some Characteristics of Water. Water (H_2O) when pure is colorless in small amounts, tasteless, and odorless, and in this state it is a non-conductor of electricity. The water (H_2O) of oceans, lakes, and rivers has a blue or green color, and this is due to the natural color of the gases of which it is formed, the refraction and reflection of the light that strikes it, and to the mineral and other substances in it.

Like other liquids, water (H_2O) is almost incompressible and it remains a liquid at temperatures between 32 degrees and 212 degrees of *Fahrenheit's* thermometer. At 32 degrees it freezes into a solid which we call *ice,* and at 212 degrees it boils and passes into the air as a vapor which we call *steam.* To reach this form it expands 1700 times in volume, or bulk, which means that 1 pint of water will make 1700 pints of steam.

What Water is Made of. Water (H_2O) is formed of two elements and these are hydrogen (H) and oxygen (O), and, as you know, both of these are gases. To form water (H_2O) they must be chemically combined in the proportions of 2 parts of hydrogen (H) to 1 part of oxygen (O), that is, $\mathrm{H_2O}$, by volume, or bulk, or in the proportion of 2 parts of hydrogen (H) and 16 parts of oxygen (O) by weight, which is the same thing. These measures are easily proved to be correct both by analyzing, that is, decomposing, water (H_2O), and by measuring and weighing the gases separately, and also by taking these gases in the above proportions and chemically combining them, upon which synthetic[1] water (H_2O) results. How to analyze water (H_2O) and how to produce it synthetically will be explained in the next chapter.

What Water Is Good for. Water (H_2O) is not only necessary to drink, to bathe in, and for the construction of living plants and animals and their maintenance, but it has many other uses as well. For instance, it is one of the chemist's allies in that it is a great solvent, for more substances can be dissolved in it than in any other liquid, hence, it is always used first when a substance is to be analyzed and it forms the basis of many solutions. Because it cannot be com-

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[1] A synthetic compound is one that you have built up of the same elements as those used by nature.

pressed, it is used in hydraulic presses and other machinery, while in the form of ice it is largely used as a cooling medium, and in the form of steam it has a wide application as a power, or prime mover.

How to Purify Water. Water (H_2O) is never found pure in nature; rain water (H_2O) is the purest, but even that has foreign matter in it. In making chemical experiments where water (H_2O) is to be used, it must be pure or the results may not be at all what you expect them to be. Now water (H_2O) can be purified by several methods, and chief among those are by filtration, by boiling, and by distillation. Where water (H_2O) is filtered, only the larger particles of matter in it are removed. Boiling kills all of the germs in it, and much of the suspended matter will fall to the bottom when it is allowed to settle, so that for ordinary experiments you can use boiled water (H_2O), which should then be filtered. The only way to get pure water (H_2O), though, is to distil it.

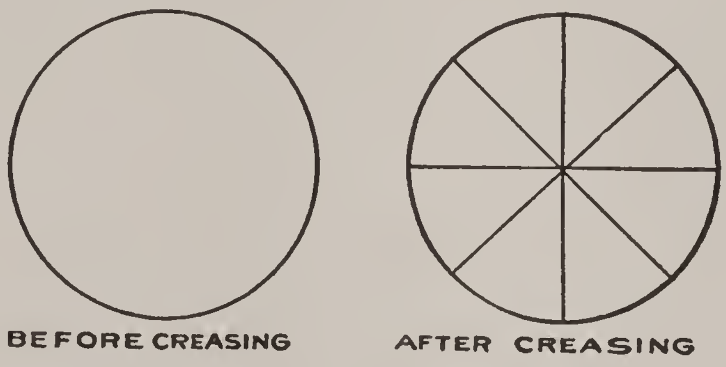

FIG. 61.—How the Filter Paper is Creased.

How to Filter Water. To filter water (H_2O) in order to get rid of any solid particles in a solution you need a glass funnel, as shown in Fig. 17 in Chapter I. The filter paper comes in round sheets and you can fit it into the funnel by folding a sheet of it over once, then again, and, finally, again, causing it to be creased, as shown in Fig. 62.

This done, spread the paper out flat and then make a cone of it, set it into the funnel and rub it along the creases to make it fit closely, as in Fig. 62; next, wet the paper all over with clean water to make it cling to the surface, and set the funnel in the ring of a support, as in Fig. 63. Finally place a beaker, or a test tube under the funnel and pour the solution you want to filter into the latter.

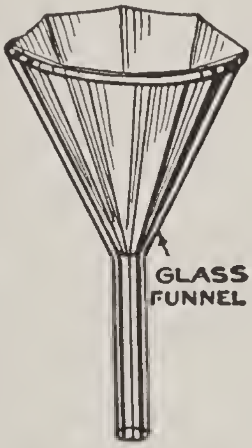

FIG. 62.—The Filter Paper in the Funnel.

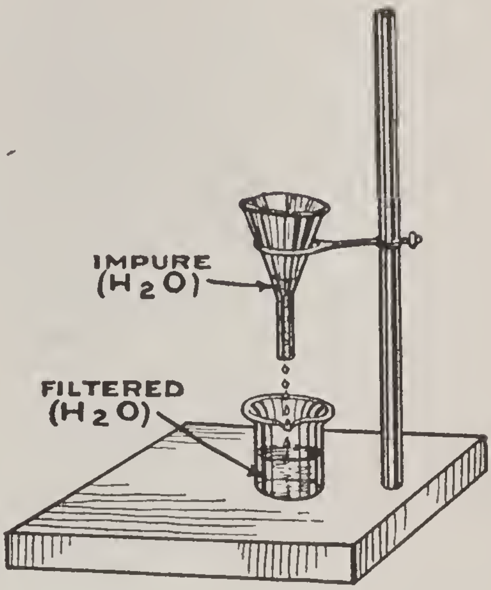

FIG. 63.—The Funnel in Use.

How to Boil Water. You can boil the water (H_2O) in an ordinary teakettle, or if you only need a small amount of it you can use a beaker. Boiling does not remove all the foreign matter in the water (H_2O) by any means, but if it has what is called *temporary hardness*, which will be explained presently, then the mineral compounds causing it will be deposited on the sides and bottom of the vessel and in this way are removed. It is these compounds that form *fur* in the kettle and *scale* in a boiler. But if the water (H_2O) has *permanent hardness*, boiling will not remove the compounds that cause it.

How to Distil Water. The easiest way to get distilled water (H_2O) for your experiments is to buy it at the drug store, but I shall give you three modifications of the same apparatus, so that you can distil it for yourself. To distil a very small quantity of water (H_2O), so that you can see clearly the exact nature of the process, all you need is two test tubes, a delivery pipe, an alcohol lamp, and a beaker, or a tumbler.

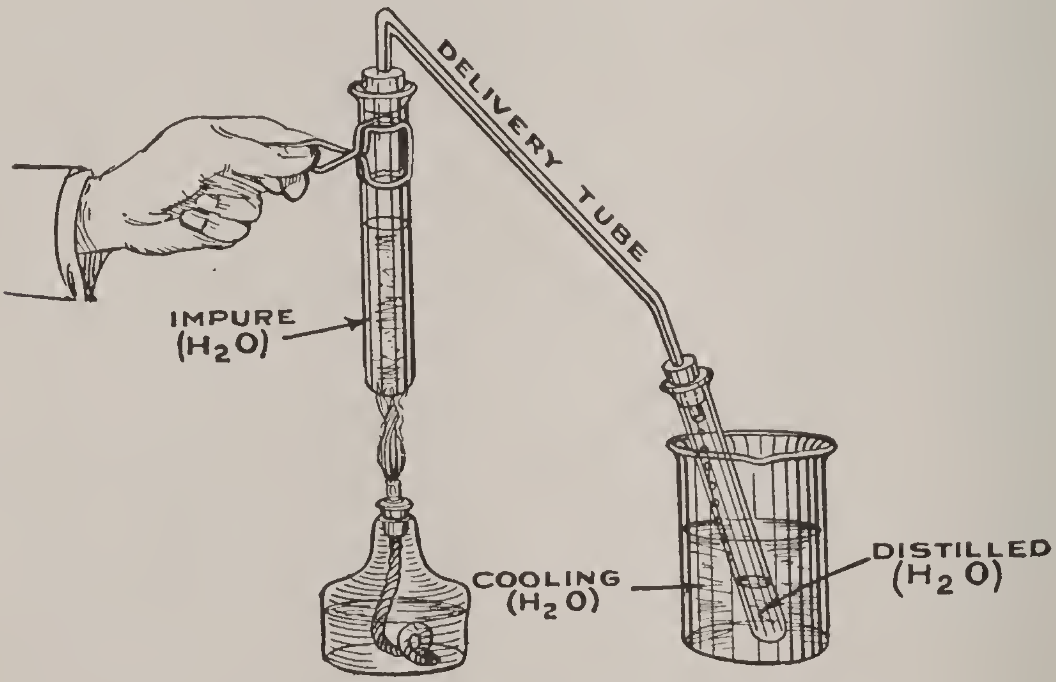

FIG. 64.—How to Distil a Little Water.

Pour enough water (H_2O) of any kind into one of the tubes to half fill it, then push the short end of the delivery tube through a cork and fit this into the neck of the test tube; put the other test tube into a beaker or tumbler of cold water (H_2O) and put the other end of the delivery tube into this second test tube. Light the alchohol lamp and hold the test tube with the water (H_2O) in it over the flame with your test-tube holder, as shown in Fig. 64, and let the water (H_2O) boil.

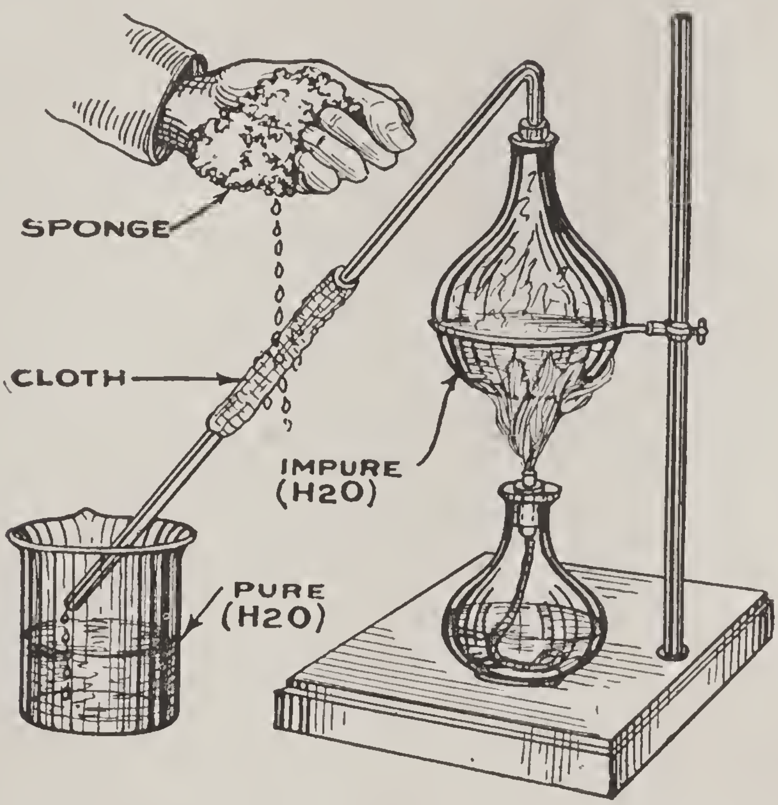

FIG. 65.—A Better Apparatus for Distilling Water.

How the Experiment Works. As soon as the water (H_2O) begins to boil, it will generate steam, and as this passes through the delivery tube it will be chilled and condensed into water (H_2O) when it reaches the cold test tube that is in the beaker.

NOTE.— To see that only pure water (H_2O) passes over and that the impurities are left behind, you can dissolve enough cupric sulphate ($CuSO_4$), or *copper sulphate*, *bluestone*, or *blue vitriol*, as it is variously called, in the water

(H_2O) you are going to distil to give it a good green color, and you will see that this is left behind.

To distil enough water (H_2O) to make an experiment with, half fill a small glass flask with some water (H_2O) and set it in the ring of your stand. Now place your lamp, or burner, directly under the flask, and put a beaker, or a tumbler, under the end of the delivery tube, as shown in Fig. 65. This done, light the lamp and when the water (H_2O) begins to boil and to generate steam, the latter will pass through the tube; when it does so, let some cold water (H_2O) fall on the tube by means of a sponge, and the steam (H_2O) will then condense into water (H_2O). The better to aid the process of *condensation*, wrap a cloth round the tube and let the water fall on it. Very soon pure water (H_2O) will flow out of the tube and into the beaker.

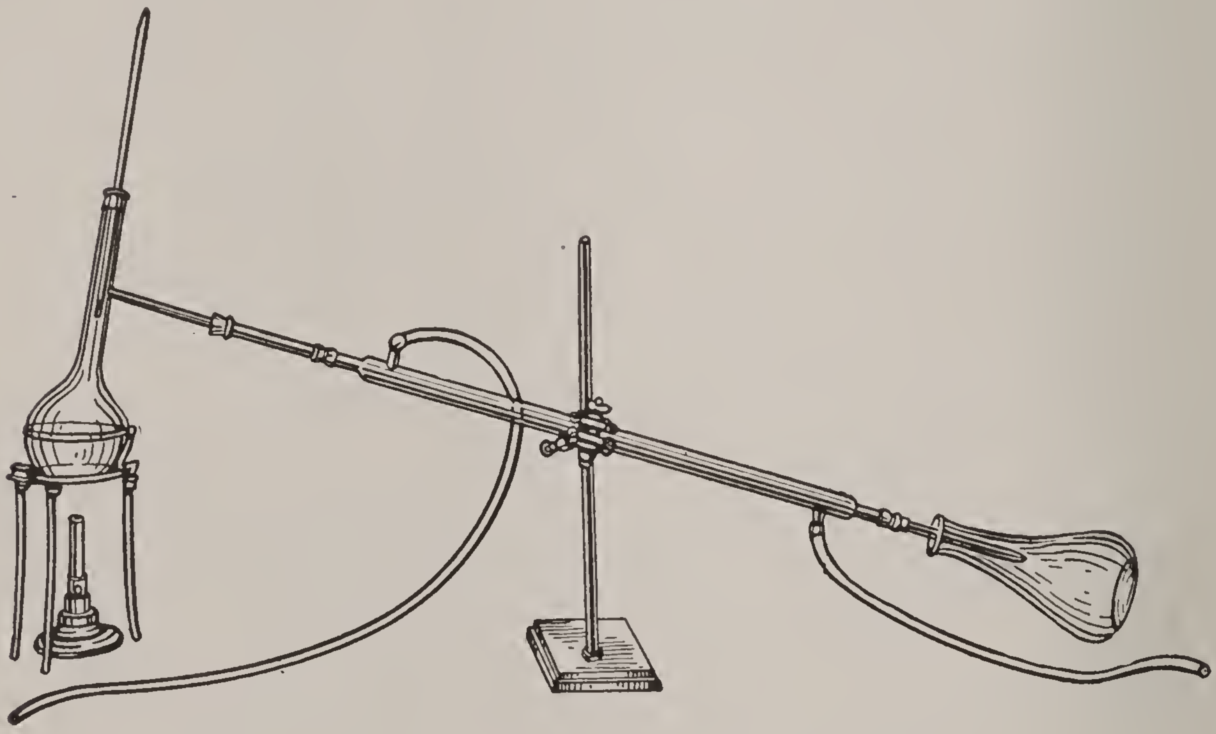

FIG. 66.—An Apparatus for Distilling Water on a Large Scale.

An apparatus for distilling water in large quantities is shown in Fig. 66. It consists of a retort, with a delivery tube which passes through a larger tube sealed to the former at both ends, so that as much of the surface of the delivery tube as possible will be exposed to the cooling water (H_2O). A stream of cold water (H_2O) is made to flow into the cooling tube at the bottom and to flow out of it at the top, as warm water (H_2O) always rises when circulating.

Tests for the Purity of Distilled Water. The first proofs of the purity of water (H_2O) are that it has no color, no odor, and no taste, and that it is perfectly clear and transparent. Farther, it must not change the colors of indicators, such as litmus paper and phenolphthalein ($C_{20}H_{14}O_4$), and, finally, when slowly evaporated it must not leave any solid matter behind. How to make a test for each of these will be explained farther along.

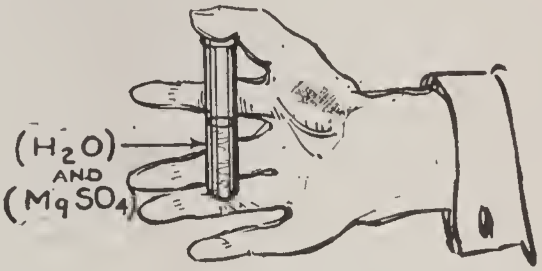

Fig. 67.—How to Raise the Temperature of Water.

How to Raise the Temperature of Water. Pour enough distilled water (H_2O) into a test tube to half fill it, then hold it by the mouth and place the closed end against your cheek, which will give you a rough idea of the temperature of the water (H_2O). Now put ½ teaspoonful of magnesium sulphate ($MgSO_4$), or *Epsom salts*, as it is commonly called, in the tube, hold your finger, or thumb, over the mouth of it, as shown in Fig. 67, and shake it until the salts have completely dissolved. Again hold the tube to your cheek and you will find that it is considerably warmer than it was before.

How the Experiment Works. Many compounds besides magnesium sulphate ($MgSO_4$) have what is called a *positive heat of solution*, and when they come in contact with water (H_2O) they give up their latent heat to it.

How to Lower the Temperature of Water. Pour enough distilled water (H_2O) into a test tube to make it half full and hold it to your cheek to get an idea of its temperature as before. This time put ½ teaspoonful of ammonium chloride (NH_4Cl), or *sal ammoniac*, as it is more often called, in the tube and shake it until the compound is completely dissolved. Again hold the tube to your cheek and you will find that it is considerably colder than it was before.

How the Experiment Works. Many compounds besides ammonium chloride (NH_4Cl) have what is known as a *negative heat of solution* when they are brought into contact with water (H_2O), and, hence, they absorb the heat of the latter.

How to Make Ice. The principle of extracting the heat of a compound by adding a substance that has a negative heat of solution is used in a practical way in making ice-cream. In this case, however, sodium chloride ($NaCl$), which is common salt, is mixed with cracked ice and this is packed around the can containing the cream to be frozen. As the ice melts and the salt dissolves, they extract the heat of the water (H_2O) thus formed, and a temperature still lower than that of the melting ice alone will be produced.

To make a little ice (H_2O), all you have to do is to pour enough water (H_2O), distilled or otherwise, into a test tube to half fill it, then put it into a beaker, or a tumbler, and fill this up with a mixture of finely cracked ice (H_2O) and soldium chloride ($NaCl$), as shown in Fig. 68. Grip the test tube by the mouth and turn it rapidly around in the beaker, and in a couple of minutes the freezing mixture will change the liquid water (H_2O) into ice (H_2O).

What Water of Crystallization Is. There are some kinds of crystals which seem to be perfectly dry, as, for instance, *Glauber's salts*, which is the *decahydrate*[1] of sodium sulphate ($Na_2SO_4.10H_2O$), that are formed of more than half of their weight of water (H_2O), and this is called *water of crystallization.* If the crystals are heated and the water (H_2O) is driven out of them, they will decompose and crumble to pieces.

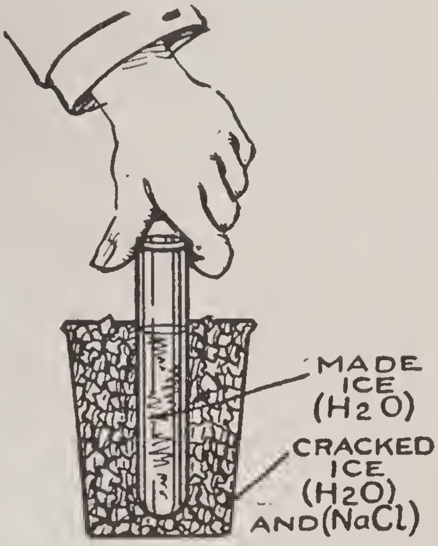

Fig. 68.—How to Make Ice.

Some crystalline compounds must be heated to the temperature of boiling water (212°F) before they will give up their water of crystallization, and others will do so when they are simply exposed to the open air. When crystals give up their water of crystallization the process is called *efflorescence.* There are, however, some crystals which when the water of crystallization is driven out of them will absorb it again when the compound of which they are formed is exposed to moist air, and new crystals are produced; this process is just the reverse of efflorescence and is called *deliquescence.*

How to See the Water of Crystallization. Put half a

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[1] A *hydrate* is a substance that combines with water, or the elements of water, and a decahydrate is a hydrate one molecule of which combines with 10 molecules of water.

teaspoonful of copper sulphate ($CuSO_4.5H_2O$) into a clean, dry test tube and heat it over the flame of your lamp or burner. Almost instantly it will boil and give up its water of crystallization in the form of steam (H_2O), and some of this will condense into little drops of water (H_2O) on the surface of the tube. You will also see that as the water is driven out of the crystals they change from blue until they become colorless. Nearly all crystals lose their colors when the water of crystallization is driven out of them.

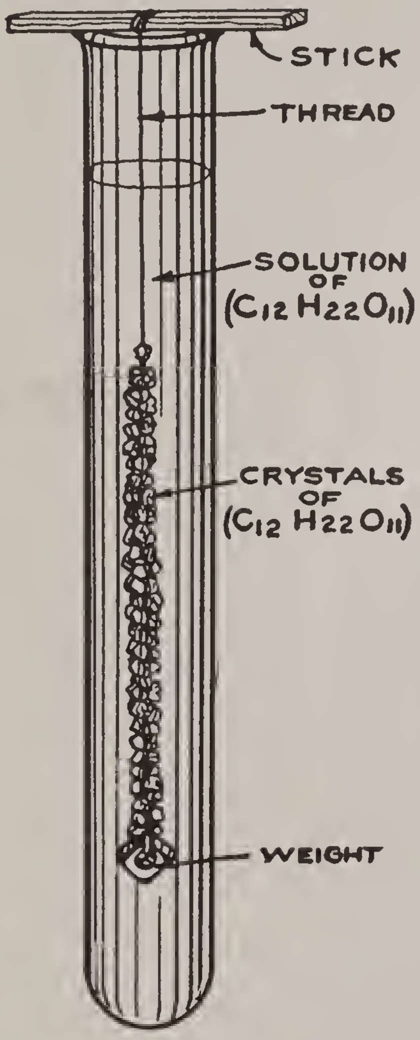

Fig. 69.—How to Make Rock-Candy.

How to Make Rock-Candy Crystals. This is an experiment in deliquescence, and to make it take a large test tube and then tie one end of a stout thread, which is about as long as the tube, to a nut, or other little weight and the other end of it to a bit of wood. This done, half fill the tube with boiling water (H_2O) and stir in as much granulated sugar ($C_{12}H_{22}O_{11}$) as it will dissolve; now let the thread down in the solution and set the tube in a rack where it can slowly cool off, and large crystals of *rock-candy* ($C_{12}H_{22}O_{11}$) will be formed on the thread, as shown in Fig. 69. In the same way you can produce beautiful crystals of other substances that have water of crystallization in them.

How the Experiment Works. When the crystals of rock-candy ($C_{12}H_{22}O_{11}$) are forming they leave behind them the water of crystallization that is in the minute crystals of sugar ($C_{12}H_{22}O_{11}$). It is the water of crystallization that makes the crystals of ordinary sugar ($C_{12}H_{22}O_{11}$) as soft as they are, and since there is very little water (H_2O) in the crystals formed on the string, they are quite hard, hence the name rock-candy ($C_{12}H_{22}O_{11}$).

How to Make a Secret Writing Ink. Put a little water (H_2O) into a test tube and add as much cobalt chloride ($CoCl_2$) as it will dissolve, after which it is called a *saturated solution.* To help along the operation, put your thumb over the mouth of the test tube, as shown in Fig. 67, and shake it vigorously. Now take a quill pen, or sharpen the end of a match, and write upon a sheet of pink paper with it. Then let it dry, and the writing will be invisible.

To read what you have written, your friend has only to heat the paper a little, and the writing will come out in a bright blue color; but as soon as the paper cools off, the writing will vanish as completely as if it had never been, at least as far as the human eye is concerned.

How the Experiment Works. Cobalt Chloride ($CoCl_2$) comes in the form of blue crystals, and these have very little, if any, water (H_2O) in them. When it is dissolved in water (H_2O) and is used as an ink, the cobalt chloride ($CoCl_2$) absorbs the moisture from the air and forms crystals that, of course, contain water of crystallization, and the mixture of cobalt chloride and water (H_2O) in them has the formula of ($CoCl_2+6H_2O$), since the cobalt chloride ($CoCl_2$) has combined with 6 molecules of water (H_2O).

How to Make a Weather Forecaster. The same principles can be used for indicating whether the weather is going to be *fair* or *rainy*, and this is done by the percentage of moisture there is in the air. Put enough water (H_2O) into a test tube to fill it half full, and then dissolve all the cobalt chloride ($CoCl_2$) that it will take up, that is, you make a saturated solution of it.

Cut a strip ½ inch wide and 4 inches long from a sheet of clean white blotting paper and immerse it in the cobalt chloride ($CoCl_2$) solution; now hang it up to dry and it will forecast what the weather is to be by its changing colors. When *rain* is to be expected, the air will be *damp* and the moisture will turn the crystals of cobalt chloride ($CoCl_2$) in and on the paper *pink*, and, when the weather is to be *fair*, the air is much *drier* and the crystals lose enough of their water of crystallization to turn them *blue*.

How to Make Imitation Ground Glass. A sheet of clean glass painted over with the following solution and then allowed to dry makes a very good imitation of *ground glass*. To make the solution, nearly fill a large test tube with water (H_2O) and then put in 3 teaspoonfuls of ammonium chloride (NH_4Cl), which is sal ammoniac; shake the tube until the salt is thoroughly dissolved, then stir in a couple of drops of glue, boil it over the flame of your alcohol lamp, and paint the surface of the glass with it while it is hot. As soon as the solution begins to cool, the water (H_2O) will start to evaporate and minute crystals will form all over the surface of the glass, and it will look as if it were ground.

Kinds of Water. By kinds of water (H_2O), are meant various specimens of it that contain different substances, but, of course, in the last analysis there is really only one kind, and this is water (H_2O) — that is the liquid formed by the chemical union of 2 parts of hydrogen (H) and 1 part of oxygen (O) by volume.

Water (H_2O) is classified under two general heads, and these are *soft* water and *hard* water. When water (H_2O) is pure, or nearly so, it is called soft water, and when it contains mineral substances it is called hard water. Of hard water there are also two kinds, namely, those that have *temporary hardness* and those that have *permanent hardness.*

How to Tell if Water Is Soft or Hard. You can easily find out whether water (H_2O) is soft or hard by rubbing up some soap with it. If it lathers well it is soft water, and it follows that if it does not lather well it is hard water.

How to Test For and Get Rid of Temporary Hardness. Half fill a test tube with some of the water (H_2O) to be tested and boil it for several minutes over the flame of your alcohol lamp. If it contains calcium carbonate ($CaCO_3$), or limestone, as it is commonly called, it has temporary hardness, and to get rid of it you need only to boil it, upon which the limestone ($CaCO_3$) will be precipitated, that is, thrown down to the bottom of the tube; in this way boiling the water (H_2O) makes it soft.

How the Experiment Works. When rain falls it absorbs the carbon dioxide (CO_2) in the air, and as the water (H_2O) containing the gas filters through the earth, the carbon dioxide (CO_2) acts on the limestone, or calcium carbonate ($CaCO_3$), and changes it into the more soluble form of calcium bicarbonate ($CaCO_2$), which is dissolved by and remains in the water (H_2O). Now when the water is boiled, the carbon dioxide (CO_2) is driven off and the calcium bicarbonate ($CaCO_2$) again becomes ordinary limestone, or calcium carbonate ($CaCO_3$), and this is precipitated, that is, it is thrown down to the bottom of the vessel.

How to Test For and get Rid of Permanent Hardness. Half fill a test tube with some of the water (H_2O) to be tested and then add ½ teaspoonful of sodium carbonate ($Na_2CO_3+10H_2O$), which is commonly called soda; place your thumb over the mouth of the tube and shake it hard. If, now, the water has permanent hardness, white particles of *gypsum*, that is, calcium sulphate ($CaSO_4$), or *Epsom salts*, which is magnesium sulphate ($MgSO_4$), or both of these salts, will be precipitated, mainly in the form of carbonates. Then run the water through a sheet of filter paper, and the particles will be left behind and the water will be soft.

How the Experiment Works. When sodium carbonate ($Na_2CO_3+10H_2O$) comes in contact with sulphates that cause permanent hardness, the latter are decomposed and form calcium carbonate ($CaCO_3$) and magnesium carbonate ($MgCO_3$), which then fall to the bottom, thus leaving the water (H_2O) soft. How soap cleans when it is used with water (H_2O) is explained in Chapter XV.

How to Test Water for Odor and Color. Half fill a test tube with some of the water (H_2O) to be tested, shake it well, and then hold it to your nose; if the water (H_2O) contains any living organisms, it will give off an odor, and if there are any decaying organic impurities in it, the odor may be an unpleasant one. Now heat the test tube and again smell of it, when you may find that the odor is even more pronounced.

Half fill a test tube with water (H_2O) and let it stand for a few minutes; now hold if between your eyes and a sheet of white paper against the light, and you can easily see if it is tinted or not and whether it is clear or translucent. This done, shake it well and then examine it again.

How to Test Water for Mineral Substances. Pour a little of the water (H_2O) to be tested into a watch crystal, or, better, a small, porcelain evaporating-dish, and set it in the sun until all the water has evaporated. If you want to evaporate it more quickly, heat it very gently over the flame of your lamp or burner, as shown in Fig. 70. If it contains organic, or mineral matter of any kind, or both, such matter will be left behind as a *residue* on the crystal, or dish. Now heat it to redness, and if the residue is formed of organic matter it will be decomposed and pass off in the form of gases, while if it is formed of mineral matter, it will not be affected by the heat.

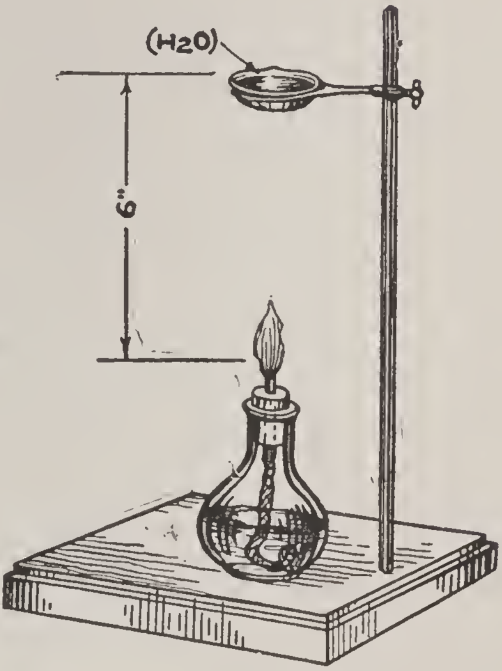

Fig. 70.—How to Test for Mineral Matter in Water.

How to Test Water for Organic Matter. By *organic matter* is meant matter that is living or was once alive. Another and more showy test for organic matter than the one given above is to fill a beaker, or a tumbler, with some water (H_2O); now put 1 teaspoonful of sodium bisulphate ($NaHSO_4$) in a test tube half full of water (H_2O) and shake it until it is dissolved. This done, put 10 or 12 drops of the solution into the water (H_2O) to be tested.

Now dissolve ¼ teaspoonful of potassium permanganate ($KMnO_4$) in a test tube of water (H_2O), and with your medicine dropper, or pipette, as it is more properly called, add the solution a drop at a time to the water (H_2O) you are testing, and at the same time stir it with a clean glass rod or tube, as shown in Fig. 71, until it turns a violet color. Let the beaker stand for half an hour or more, and if the color of the water (H_2O) does not change in that time you may safely conclude that there is no organic matter in it.

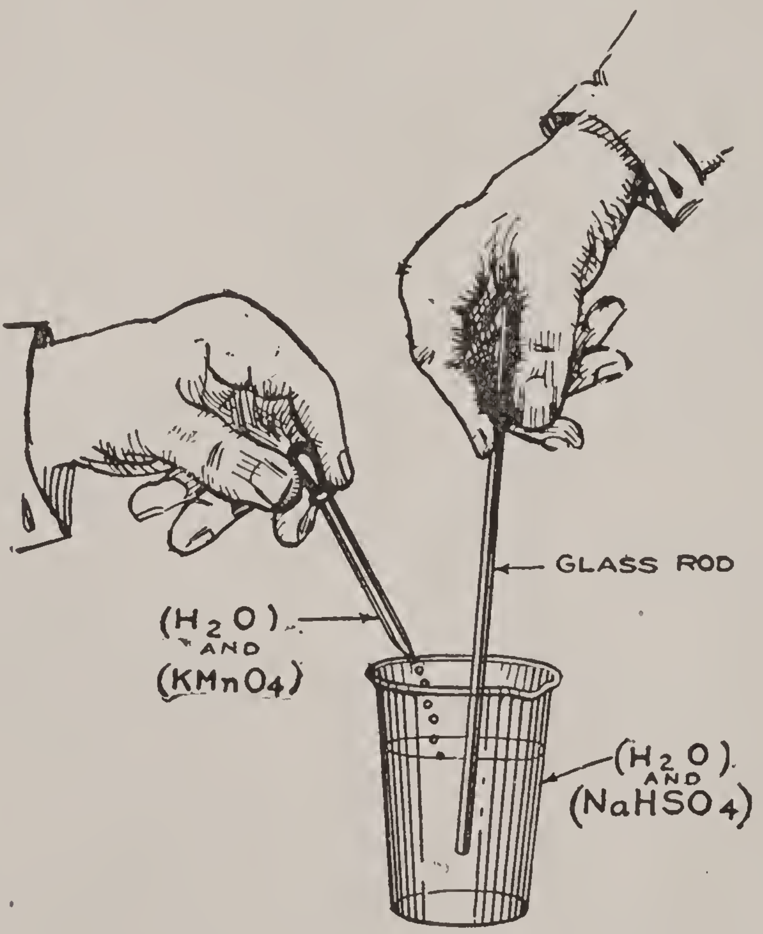

Fig. 71.—How to Test for Organic Matter in Water.

If, however, the water (H_2O) loses part of its color, it shows that there is organic matter in it.

How to Test Water for Carbon Dioxide. First prepare a little *lime-water*, which is calcium hydroxide ($Ca\,(OH)_2$), and this you do by half filling a test tube with pure water (H_2O), and then dissolving ¼ teaspoonful of quicklime, which is calcium oxide (CaO) in it. Let it settle and pour off the clear part, which is lime-water ($Ca(OH)_2$). Now fill a test tube nearly full of the water (H_2O) you want to test for carbon dioxide (CO_2) and then put in half a dozen drops of the lime-water ($Ca(OH)_2$). If there is carbon dioxide (CO_2) in the water, it will promptly take on a milky color.

How to Test Water for Alkalis. Take a test tube full of the water (H_2O) you want to test and put in a couple of drops of a solution of phenolphthalein ($C_{20}H_{14}O_4$), and this you can make by dissolving a little of it in pure methyl alcohol (CH_3OH), which goes by the name of *wood-spirit*, or *wood alcohol*. The phenolphthalein ($C_{20}H_{14}O_4$) is a colorless compound, but on coming in contact with an alkali it takes on a red tint and so colors the water.

How to Test Water for Lime. Add ¼ teaspoonful of sodium carbonate (Na_2CO_3) to a test tube of the water (H_2O) you want to test, and let it stand for half an hour. If there is no lime, that is, calcium carbonate ($CaCO_3$), in it, the water (H_2O) will remain clear, but if there is any lime in it, the water (H_2O) will take on a milky color.

How to Test Water for Acids. To make this test all you need to do is to soak a strip of blue litmus paper in the water (H_2O), upon which it will change color if it contains an acid.

How to Test Water for Iron. To a test tube that is half full of the water (H_2O) you are going to test, add ¼ teaspoonful of sodium ferrocyanide ($Na_4Fe(CN)_6+12H_2O$) and shake it well. Let the water (H_2O) stand for a few minutes, and if it takes on a blue color it shows that there is iron (Fe) in it.

How to Test Water for Sulphur. Fill a test tube full of the water to be tested, and then soak a strip of *sulphide test paper* in it. If the water contains sulphur (S), the paper will change its color to a brownish-black. *Sulphur water* is a mineral water (H_2O) that has a gas in it called hydrogen sulphide (H_2S), and it is this gas that makes it smell like rotten eggs.

CHAPTER V.

EXPERIMENTS WITH HYDROGEN

It was the English chemist, Cavendish, who first showed that hydrogen (H) was a gas by itself and, also, that it would produce water (H_2O) when it was burned in air. It was known, however, before he made these experiments that it was the oxygen (O) of the air which supports combustion, and this showed that water (H_2O) was formed of oxygen (O) and hydrogen (H) chemically combined.

Hydrogen (H) is the lightest gas known, and it is about 14½ times lighter than air, for which reason it is used for filling balloons. It is a colorless, tasteless, and odorless gas, and does not change blue litmus paper red, which shows that it has no acid properties, and yet it is a necessary element of all acid compounds. Hydrogen (H) like nitrogen (N) is not a poisonous gas, but it cannot support either combustion or life; unlike the latter gas, it burns with an intense heat in air, or better, in pure oxygen (O), and when these two gases are mixed with each other (not chemically combined) they form a very explosive mixture.

How to Analyze Water. After you have made the experiments described in the foregoing chapter your next step is to analyze some water (H_2O), that is, separate it into its two original gases, namely oxygen (O) and hydrogen (H); and then you want to make some of the latter gas and do the experiments which follow.

You can easily separate the oxygen (O) and hydrogen (H) of which water (H_2O) is formed by a process known as electrolysis—that is, by passing an electric current through it. To do the experiment you will need a pair of test tubes, a couple of pieces of carbon such as is used for arc lights and each of which is about 1½ inches long, a soup-plate, and a battery of 5 or 6 dry cells. Take two pieces of insulated copper wire and scrape the ends clean, then twist one end of each one around each of the pieces of carbon and connect the other ends to the battery of dry cells.

This done, fill the soup-plate nearly full of clean water (H_2O) — it does not have to be distilled — and stir half a dozen drops of sulphuric acid (H_2SO_4) into it. You will remember I told you in the chapter before this that water (H_2O) is not a conductor of electricity, but you can make it so — it is then called an electrolyte — by adding a little common salt, that is, sodium chloride ($NaCl$), or, better, sodium bisulphate ($NaHSO_4$) or, still better, sulphuric acid (H_2SO_4).

Now fill both test tubes full of the electrolyte, which is the water (H_2O) so prepared, then place your finger over the mouth of each one in turn, invert it and set it into the water (H_2O) in the soup-plate over the carbon rod, or electrode, as it is called, as shown in Fig. 72. As soon as you have done this you will see bubbles of gas form on each carbon electrode and rise up through the water (H_2O) to the surface of it.

Now the gas formed in one of the tubes is oxygen (O), and in the other one hydrogen (H); after this action has taken place for a few minutes you will observe that the water (H_2O) is sinking in the tubes, and after a longer interval you will further observe that there is twice as much gas in one of the tubes as there is in the other one; this is easily accounted for, since water (H_2O) is formed of 2 parts of hydrogen (H) and 1 part of oxygen (O) by volume. The gas which takes up the smaller space must, therefore, be the oxygen (O), and the one that takes up the larger space must be the hydrogen (H).

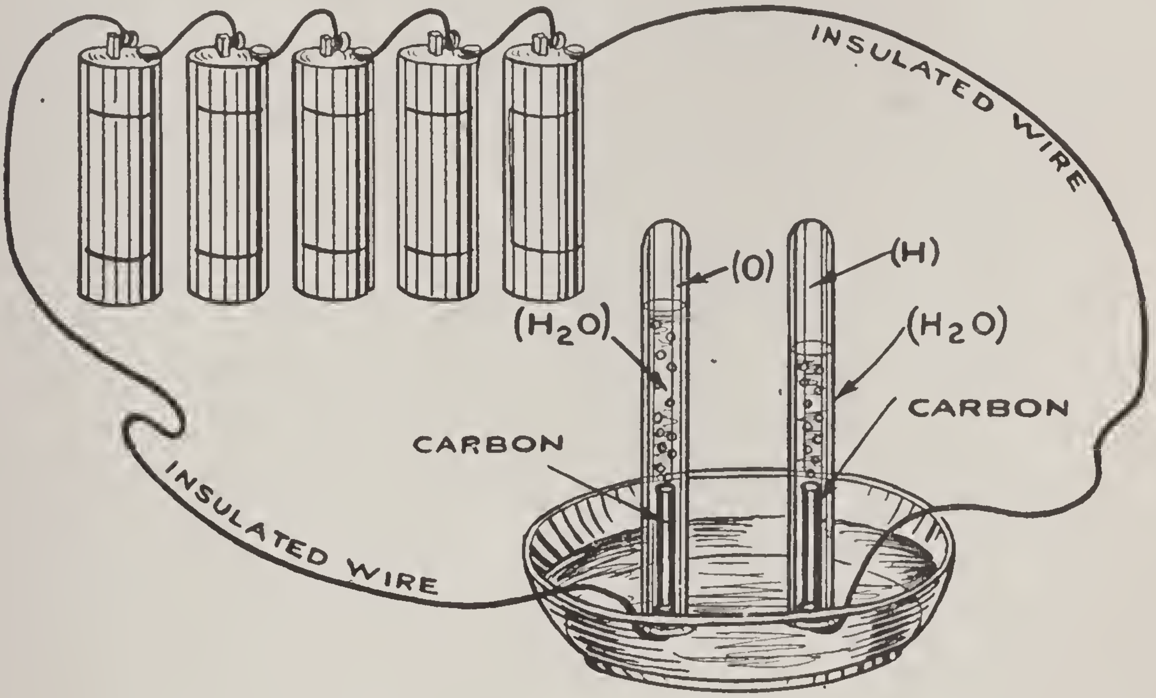

FIG. 72.—Separating Water into Its Original Gases.

To prove that these gases have really been formed in the tubes, lift up the one that has the greater amount of water (H_2O) in it and hold it mouth down, so that the water (H_2O) will run out and the oxygen (O) stay in. Now light a match, and after it gets to blazing well, blow it out and hold it in the tube; instantly it will burst into a flame again, and this shows that the gas is oxygen (O). This done, lift up the other tube and let the water (H_2O) run out of it;

next, light a match and hold it to the mouth of the tube, upon which the gas in it will explode, and this shows that it is hydrogen (*H*).

How the Experiment Works. The above experiment shows clearly enough that when an electric current acts on water (H_2O) it separates it into the two gases of which it is formed, and this process is called *electrolysis.* The discoveries made by chemists of the action that takes place when water (H_2O) and other substances in solution are decomposed by an electric current is explained on the basis of what is called *ionization.*

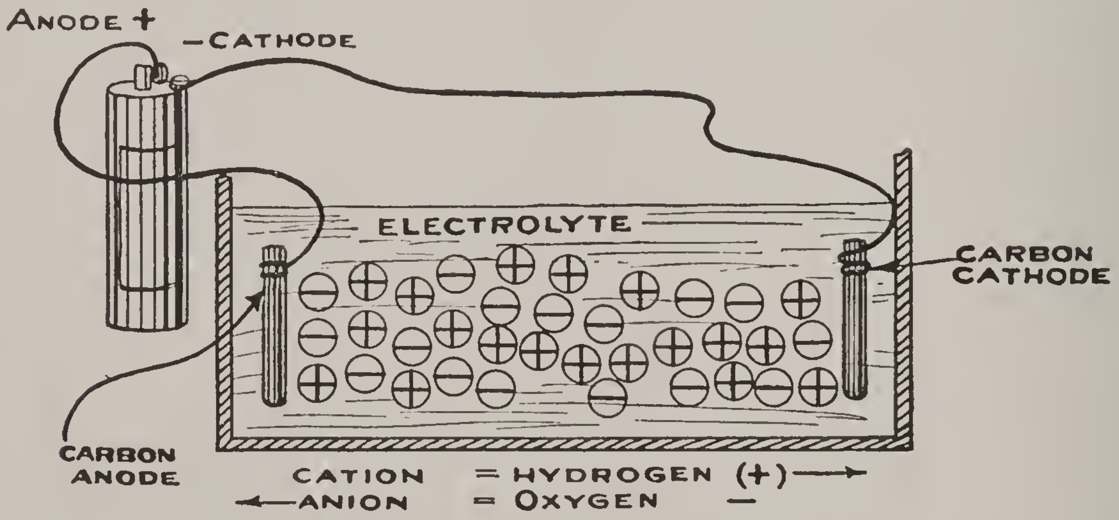

FIG. 73.—Diagram of the Theory of Ionization.

The theory of ionization supposes that the molecules which form the water (H_2O) are made to fall apart, or *dissociate*, as it is called, when an electric current flows through, and the atoms of hydrogen (*H*) and oxygen (*O*) are then *ionized* by the electric current, that is, each one takes on a charge of electricity and so they are called *ions.*

The *positively charged* atoms, or *positive* ions, are called *cations*, and the *negatively charged* atoms, or *negative* ions, are called *anions.* Now the hydrogen (H) atoms are always charged positively, and these ions of the water (H_2O), or other electrolyte, collect at the *negative* electrode, which is the carbon connected with the zinc pole of the battery, while the oxygen (O) atoms are always charged negatively, and these ions collect at the *positive* electrode, which is the carbon connected with the carbon pole of the battery, as shown in Fig. 73.

How to Make Synthetic Water with an Electric Spark. Since oxygen (O) and hydrogen (H) are obtained when water (H_2O) is decomposed, these two gases should form water (H_2O) when they are chemically combined, and this they do. The apparatus for this experiment is rather costly and is not altogether easy to make, but as it proves that oxygen (O) and hydrogen (H) when they combine form water (H_2O), I will tell you how to do it.

Fig. 74.—The Eudiometer.

First, you need a piece of apparatus called a eudiometer, and this consists of a long glass test tube; a pair of platinum wires are sealed in the wall of the tube near the closed end and form a spark gap, as shown in Fig. 74. The outside ends of the wires are connected with a small induction coil, or *spark coil,* as it is usually called, and this is energized by a battery, as shown in Fig. 75.

Now fill the eudiometer full of mercury (Hg), so that there will be no air in it, then invert it in a bowl of mercury (Hg) and keep it in an upright position with the aid of the ring-stand, as in Fig. 76. The next step is to place the free end of the delivery tube of your oxygen generator, which is shown in Fig. 42, under the mercury (Hg) and in the mouth of the eudiometer, and pass enough oxygen (O) into it to displace about 1 inch of the mercury (Hg). This done, withdraw the tube and insert one that is connected with your hydrogen generator, which is shown in Fig. 72, and pass enough hydrogen (H) into it to displace 2 inches more of mercury.

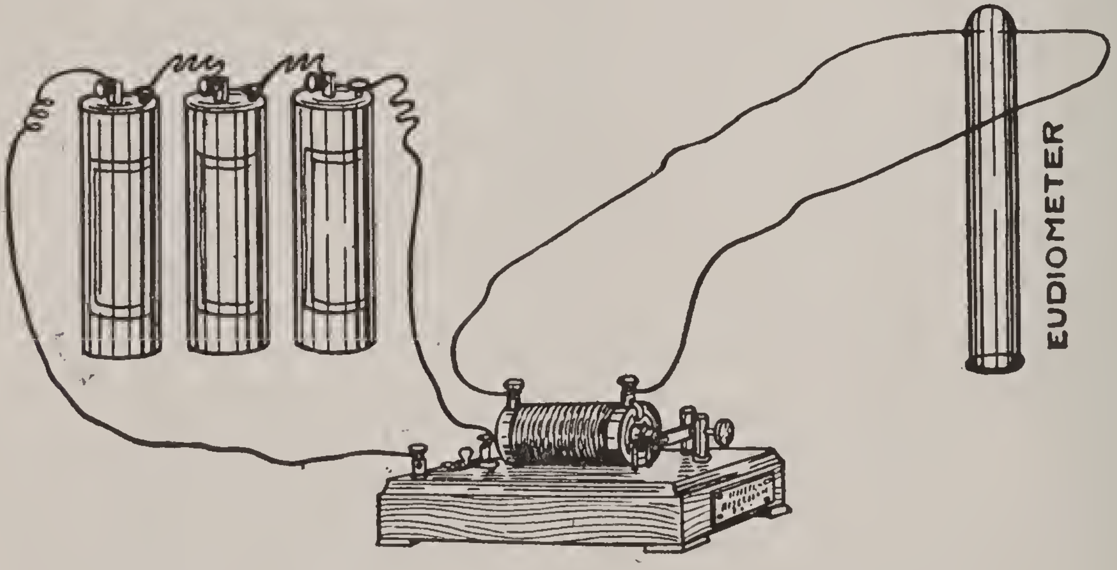

FIG. 75.—The Eudiometer Connected with the Spark Coil.

Now as long as these two gases in the tube are merely *mixed* they will remain in this condition for a long time, but the moment a spark is made to pass between the points of the platinum wires in the eudiometer, it will ignite them, a little explosion will take place, and they will combine chemically and form a minute quantity of water (H_2O).

How to Make Synthetic Water with an Alcohol Flame.

You do not need the elaborate apparatus just described to produce water (H_2O) synthetically; instead here is a very simple way in which you can generate hydrogen (H) and make it combine with the oxygen (O) of the air, and form water (H_2O).

Put a little methyl alcohol (CH_3OH), or wood alcohol, as it is called, into your evaporating-dish and light it. Now hold a perfectly dry cold beaker over the flame, and very soon minute drops of water (H_2O) will form on the inside surface of it, as shown in Fig. 77.

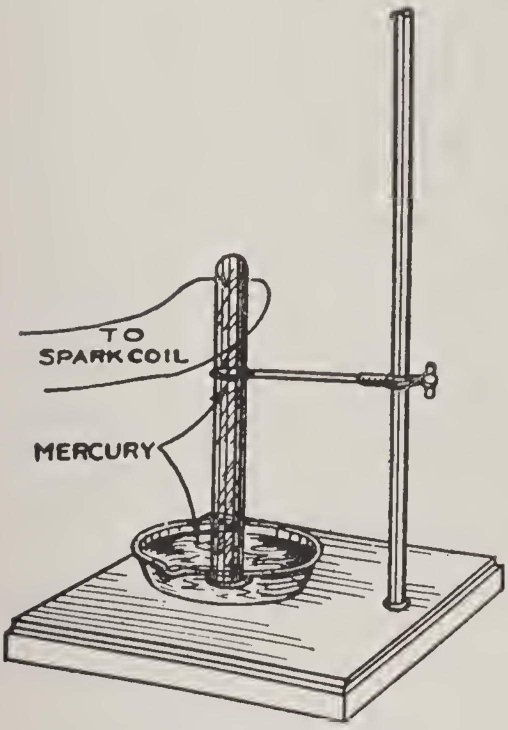

FIG. 76.—The Eudiometer Ready for the Experiment.

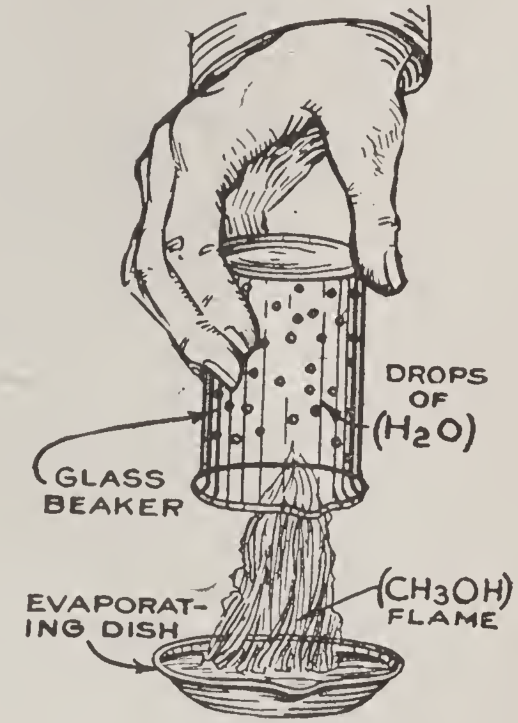

FIG. 77.—Producing Water with an Alcohol Flame.

How the Experiment Works. The alcohol (CH_3OH) contains, as the formula just given shows, 3 atoms of hydrogen (H) and the heat of the flame makes the oxygen (O) of the air combine with it, so that water (H_2O) is formed.

How to Make Hydrogen. This is the usual way that hydrogen (*H*) is made for experimental purposes. First, cut up a piece of sheet zinc (*Zn*) into bits, or better, get some granulated zinc (*Zn*) and put the zinc into an Erlenmeyer flask; now seal a glass delivery tube and a funnel tube, commonly known as a "thistle tube" from its shape, in a cork with sealing wax and put this into the mouth of the flask. Make it tight, or the hydrogen (*H*) will leak out.

This done, pour a little sulphuric acid (H_2SO_4), or *oil of vitriol*, as it is sometimes called, into the flask and add 5 or 6 times its volume of water (H_2O), and the zinc (*Zn*) will instantly act on it; the solution will boil and a great deal of heat will be evolved and a large amount of hydrogen (*H*) will be set free. As hydrogen (*H*) is so much lighter than the air, it can be collected in an inverted bottle, as shown in Fig. 78, where it will displace the air and remain for some time.

How the Experiment Works. When the zinc (*Zn*) acts on the sulphuric acid (H_2SO_4) the hydrogen (*H*) of the latter is set free and the zinc (*Zn*) takes its place, forming zinc sulphate ($ZnSO_4$) which, as its formula shows, contains zinc (*Zn*), sulphur (*S*), and oxygen (*O*). The zinc sulphate ($ZnSO_4$) thus formed is dissolved in the water (H_2O) of the acid, but you can easily recover it by evaporating the solution, upon which it will remain in the dish as a white solid.

NOTE.— Whenever you make hydrogen (*H*), you should never light it until it has passed off from the generating apparatus for at least 5 minutes. This is because there is always air mixed with the first of the gas that passes off, and this forms a very explosive mixture, due to the oxygen (O) of the former.

It is also a good plan to wrap a cloth around the flask, so that if there should be an explosion the flask will not fly to pieces. Further, always make a test of the gas first, and this you can do by filling a test tube with it and lighting it; if it burns quietly, you can then safely light it as it issues from the delivery tube.

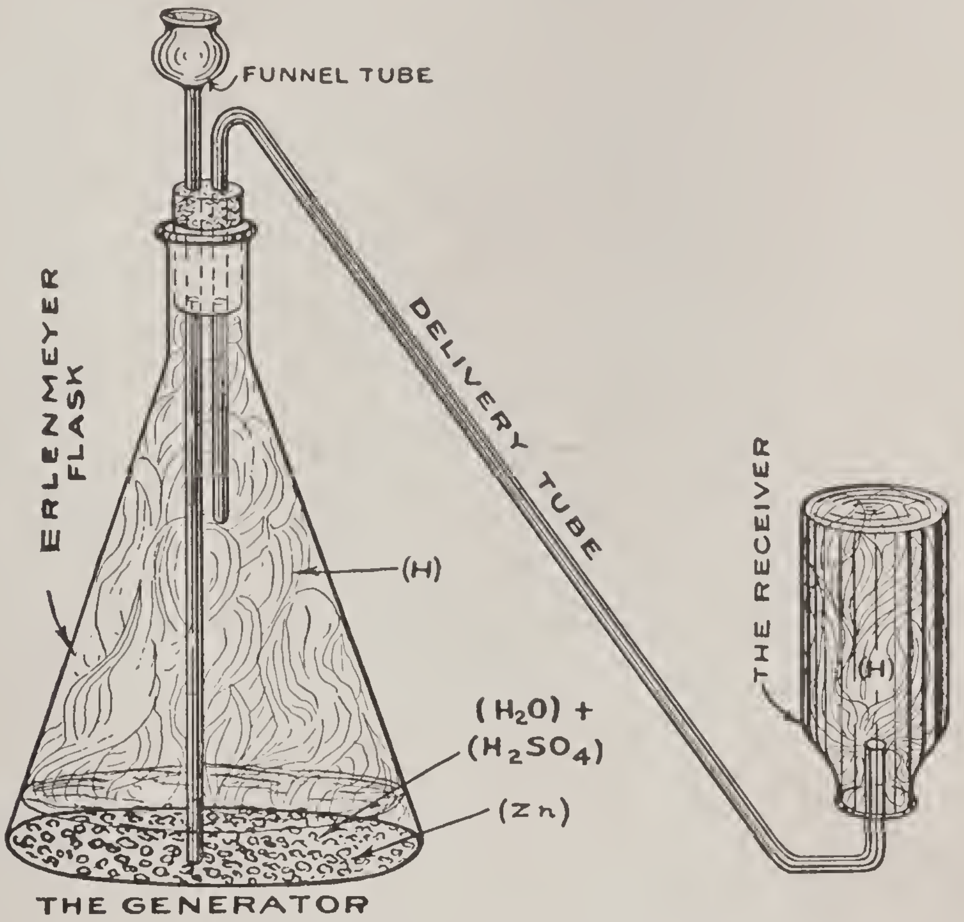

FIG. 78.—How to Make Hydrogen.

How to Make Hydrogen without an Acid. Put 1 ounce of potassium hydroxide (KOH), which is commonly called caustic potash, in an Erlenmeyer flask, or one of the ordinary kind, add ½ ounce each of fine granulated zinc (Zn) and some iron turnings (Fe) and then cover these over with water (H_2O). This done, fit a cork with a delivery tube in it into the neck of the flask; a reaction is now set up in which the hydrogen (H) is liberated, and this you can collect in another tube, or you can light it at the tip of the delivery tube.

How the Experiment Works. The zinc (Zn) acts on the potassium hydroxide (KOH) and forms potassium (K), zinc oxide (ZnO), and hydrogen (H), which is set free. By writing this reaction in the form of an equation[1] it is made clearer because of its brevity, thus:

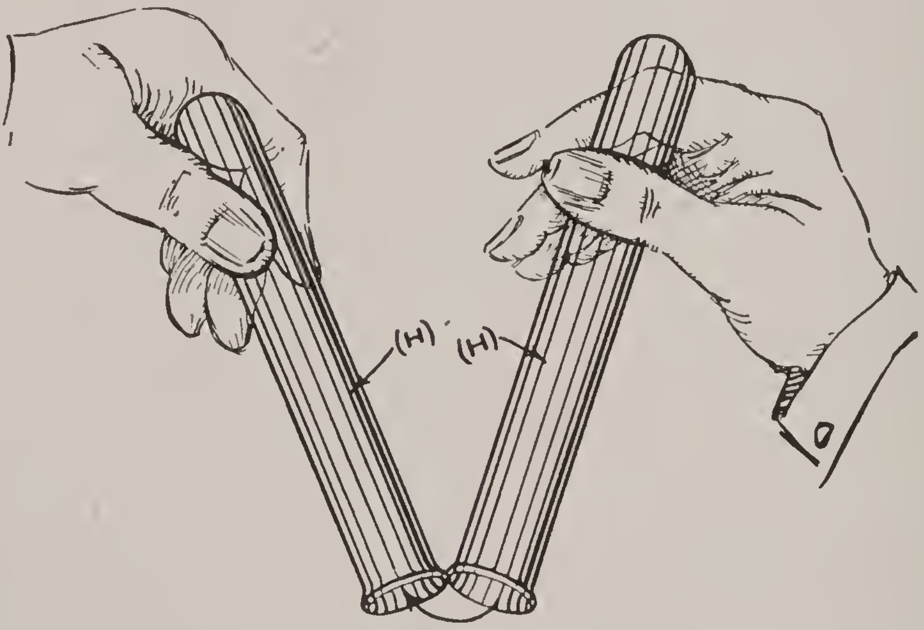

FIG. 79.—How to Pour Out Hydrogen.

Zn	+	KOH	=	K	+	ZnO	+	H ↑
Zinc		Potassium Hydroxide		Potassium		Zinc Oxide		Hydrogen

Note:—Wherever you see an arrow pointing up in an equation you will know that the preceding substance is a gas.

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[1] The nature of an equation is explained in Chapter X.

How to Pour out Hydrogen. Since hydrogen (*H*) is about 14½ times lighter than air, if you want to transfer it from one vessel to another you must pour it upward, as shown in Fig. 79. To do this, take a test tube and fill it with hydrogen (*H*), then hold another test tube vertically with its mouth down; hold the full tube vertically at first with its mouth down, with the edge of it touching the edge of the other one; now lower the closed end of the full tube and the gas will ascend and you will have performed the feat of pouring it up.

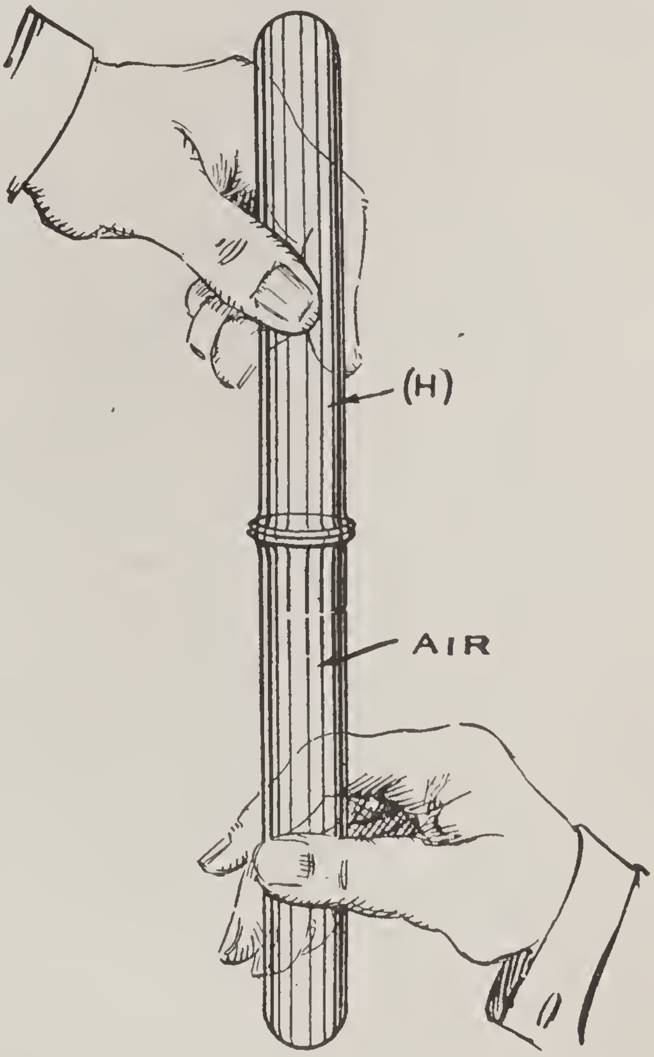

Fig. 80.—The Diffusion of Hydrogen.

The Diffusion of Hydrogen. Take two test tubes and fill one with hydrogen (*H*) and, of course, hold it with its mouth down to keep the hydrogen (*H*) in. Hold another test tube with air in it with its mouth up and place the tubes together, as shown in Fig. 80. Since air is so much heavier than hydrogen (*H*), it would seem that they would remain separated in their respective tubes, but such, however, is not at all the case; after a little while the hydrogen (*H*) sinks into the air just as though gravity were pulling it down, and this curious effect is called *diffusion.* The same action takes place when you open a bottle of perfume and its scent penetrates the air everywhere in the room.

How to Make a Hydrogen Flame. To make a hydrogen (*H*) flame, all you need to do is to take the rubber bulb from a pipette and couple the large end of it with the free end of the delivery tube of your generating apparatus by means of a bit of rubber tubing, as shown in Fig. 81. Before lighting the gas at the tip of the pipette, be sure to let the generator run at least 5 minutes to get rid of all of the air or else you are liable to have an explosion, as explained under the caption of "How to Make Hydrogen."

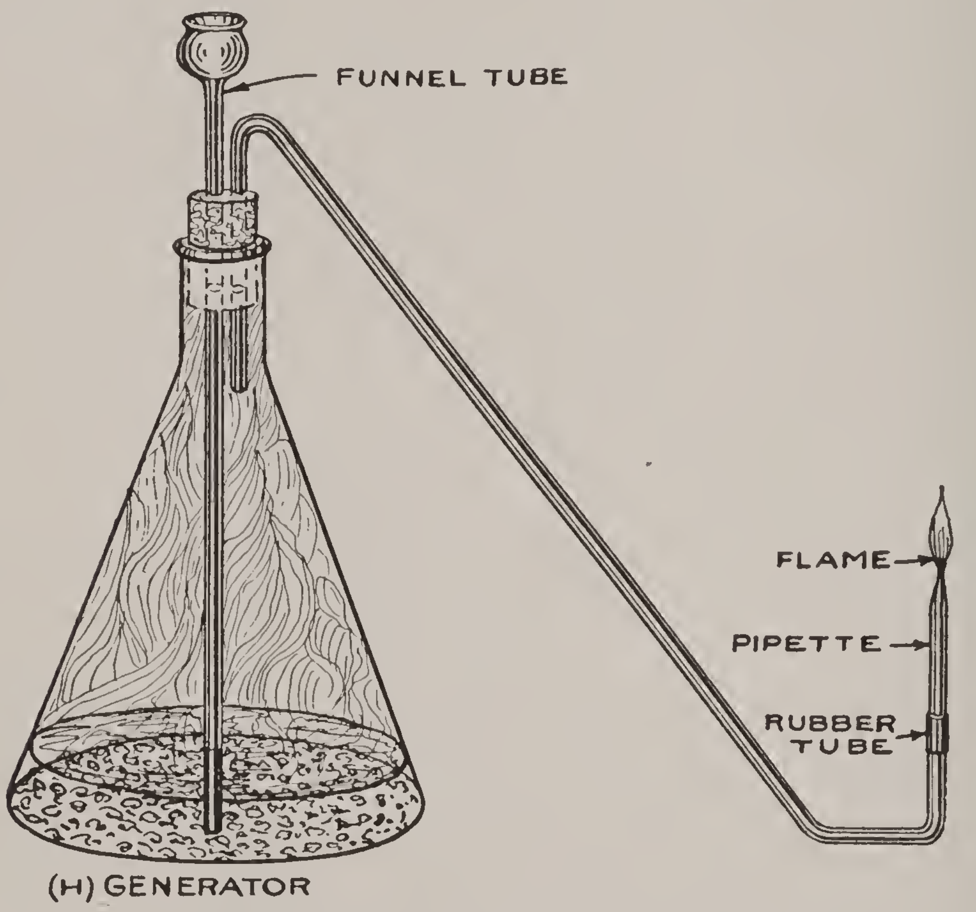

FIG. 81.—How to Make a Hydrogen Flame.

How Hydrogen Acts on Flame. Take a large test tube full of hydrogen (*H*) and keep it inverted, as shown in Fig. 82. Now light a match and hold it to the mouth of the tube and the gas will catch fire and burn with an almost invisible flame, which will work its way into the tube and finally go out. This experiment shows that hydrogen (*H*) is a combustible gas.

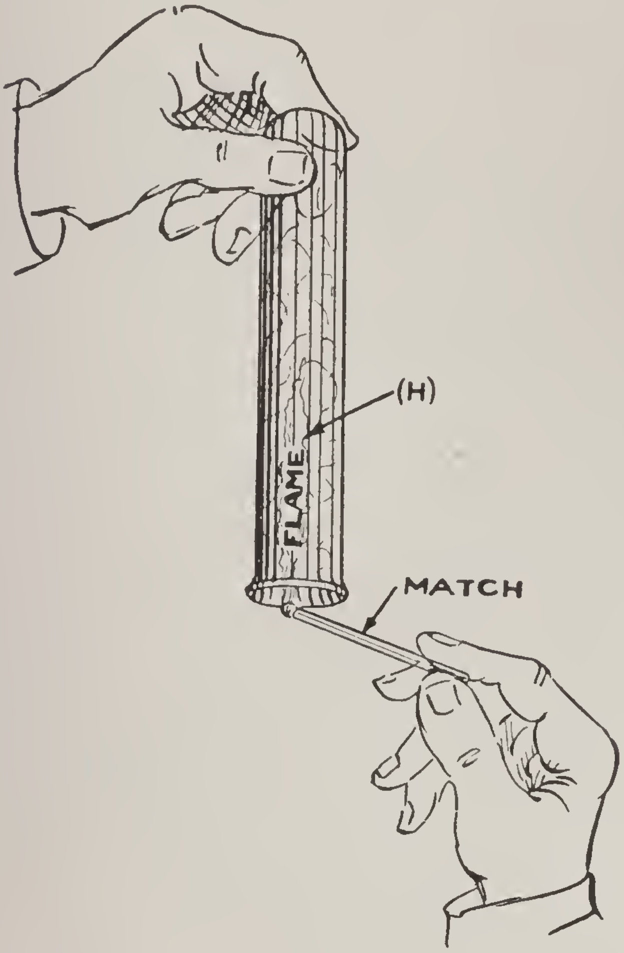

Fig. 82.—The Hydrogen Burns Gently.

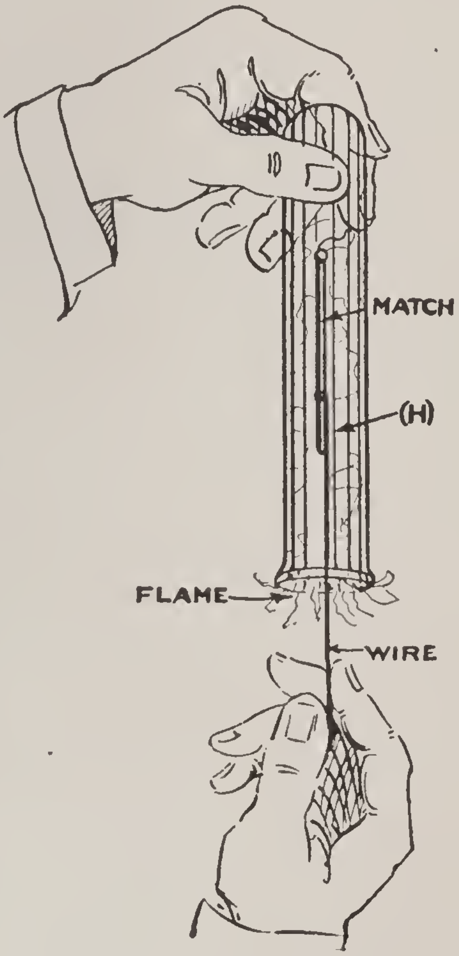

Fig. 83.—The Match is Extinguished.

Twist a wire around a match and light it. Now take another test tube of hydrogen (*H*), hold it mouth downward as before, quickly push the lighted match up to the top of the tube, as in Fig. 83, and it will go out, though the gas will burn at the mouth. This experiment shows that hydrogen (H) will not support combustion.

Finally, take another test tube full of hydrogen (H), hold it mouth upward, as in Fig. 84, and touch a lighted match to the mouth of it, and there will be an explosion. This experiment shows that when hydrogen (H) and oxygen (O) come in contact with each other they form an explosive mixture.

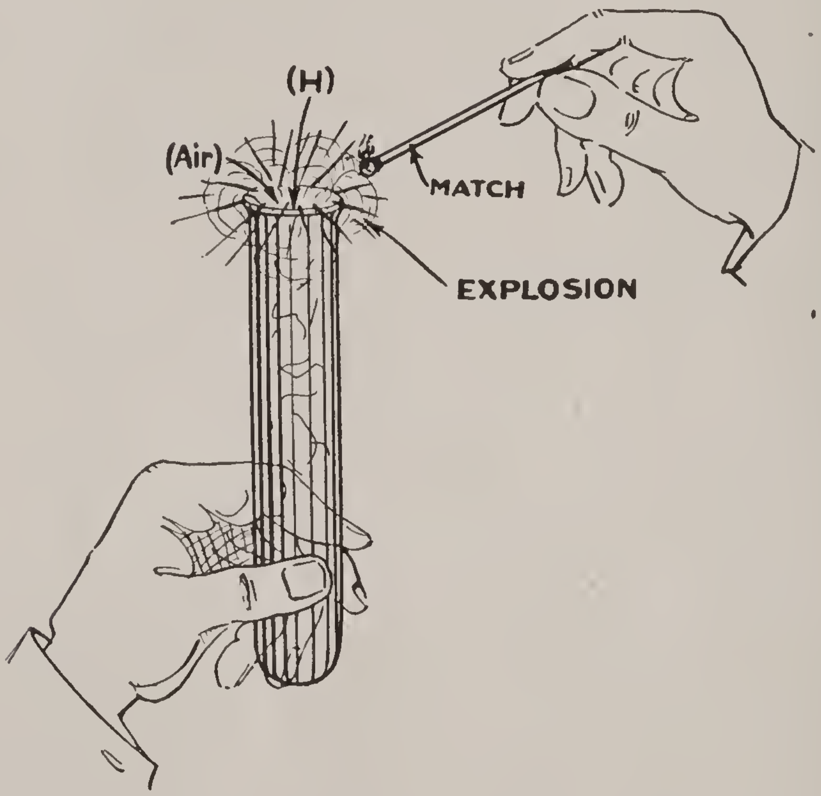

Fig. 84.—The Hydrogen Mixed with Air Explodes.

How to Blow Hydrogen Soap Bubbles. Connect a bladder[1] or, better, a small rubber gas bag, to the delivery tube of your hydrogen-generating apparatus, fill it with gas, and

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[1] To prepare a bladder for use as a gas bag, rub it well with a mixture of 1 part of glycerine and 2 parts of water. A rubber gas bag is cheap, clean, and convenient and can be bought of dealers in chemical apparatus.

then tie the neck of it up tight; this done, connect the stem of an ordinary clay pipe to the neck of the bladder, or bag, and then make a solution of good soap and soft water (H_2O).

Now dip the mouth of the pipe in the soap solution, cut the thread from the neck of the bladder, or gas bag, and then press on it, and a bubble will be formed, as shown in Fig. 85. As hydrogen (H) is so much lighter than the air, the bubble will go up like a balloon, which it really is, and break when it strikes the ceiling. If, however, you hold a lighted match to it as it ascends, it will burst with a faint yellow flash and explode with a slight noise.

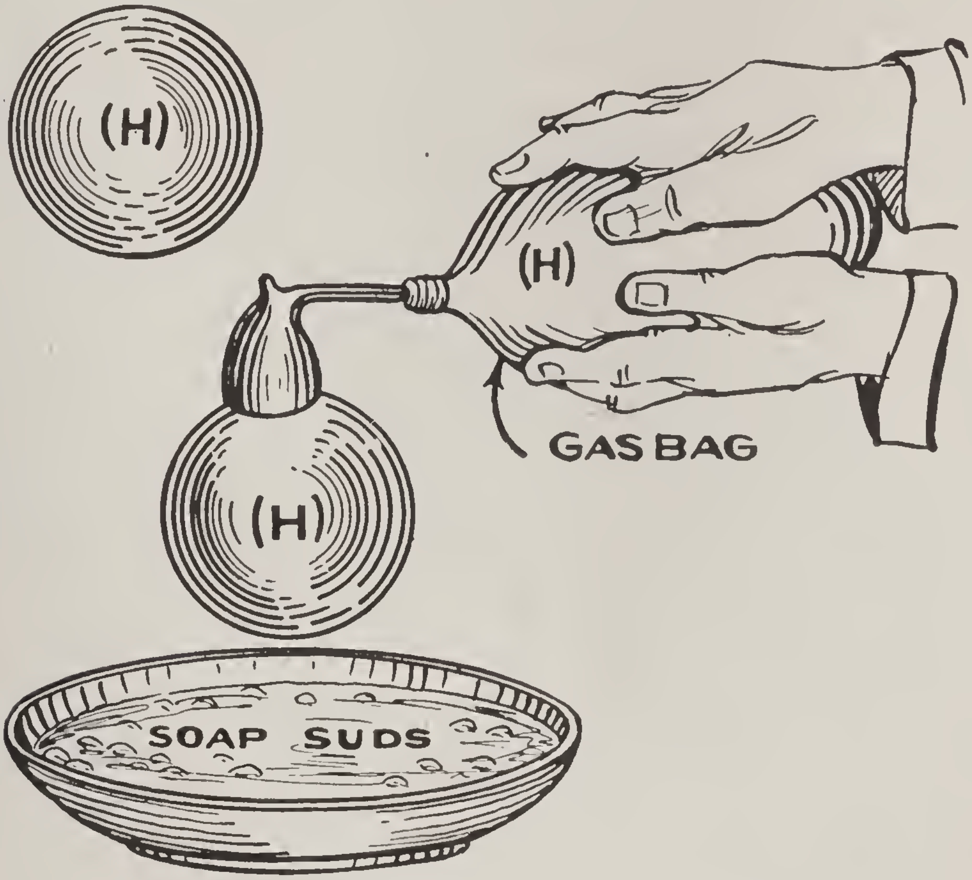

FIG. 85.—Blowing Hydrogen Soap Bubbles.

How to Blow Hydrogen Cauliflower Soap Bubbles. Fill a bladder, or a gas bag, with hydrogen (H) as before, but instead of the clay pipe fasten a glass tube in the neck of it; this done, half fill a wash-basin with the soap solution, then put the end of the tube into it and press on the bladder, or gas bag, and the basin will be filled to overflowing with small hydrogen (*H*) bubbles, as shown in Fig. 86. Now tie a match to a long stick, then light it and bring it into contact with the bubbles, and they will explode like a bunch of giant fire-crackers going off.

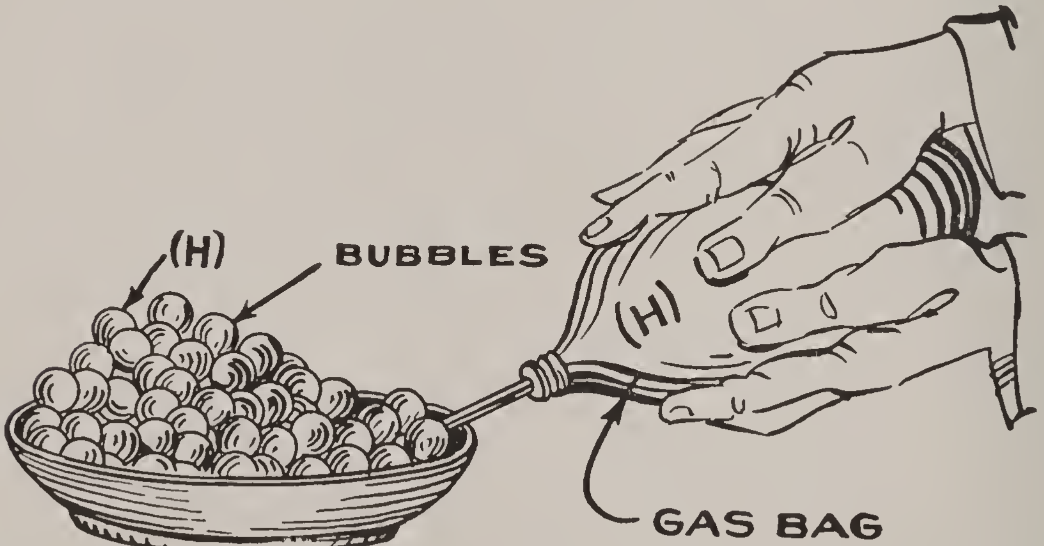

FIG. 86.—How to Blow Hydrogen Cauliflower Soap Bubbles.

How to Blow Resin Bubbles. Procure 1 ounce of pure linseed oil and 8 ounces of resin and put them in your porcelain evaporating-dish; place this dish in a pan partly filled with water. This arrangement, which is shown in cross-section in Fig. 87, is called a *water bath.* Now heat the pan with your alcohol lamp, or Bunsen burner, until the mixture is the right consistency and then blow bubbles with the clay pipe, either with air in the usual way or with the gas-bag apparatus which I have just described. If you blow resin bubbles with air, they will burst on coming in contact with the table or floor, but you can keep them for a long time by letting them fall on a sheet of paper on which you have sprinkled some lycopodium powder. Bubbles blown with the resin solution are exceedingly thin and, different from soap bubbles: they are perfectly gas-tight. They are very pretty when the sunlight is allowed to fall on them.

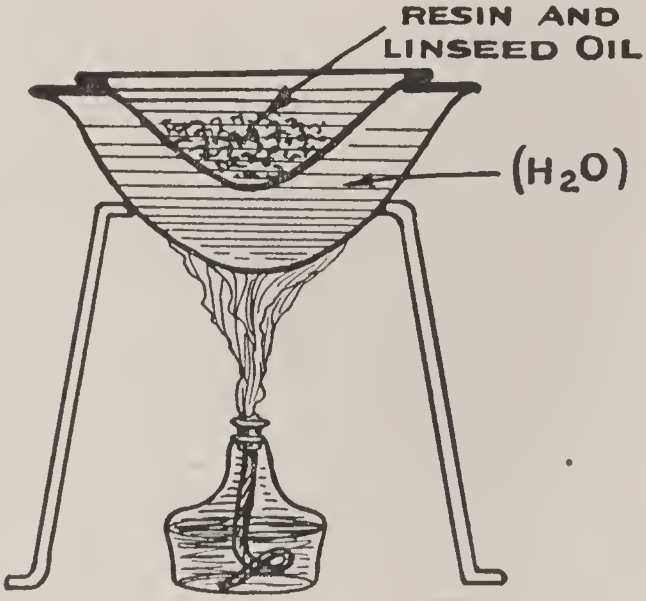

FIG. 87.—Melting the Resin and Linseed Oil Over a Water Bath.

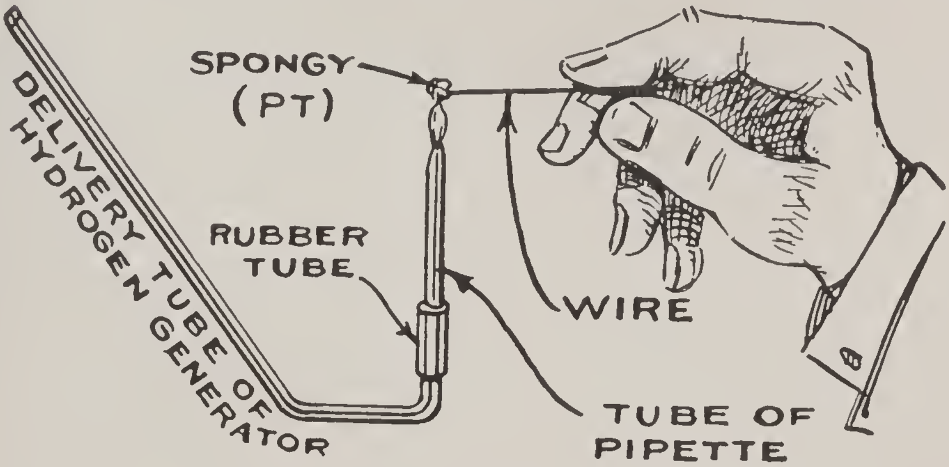

FIG. 88.—A Self-Lighting Gas Flame.

How to Make a Self-Lighting Flame. Make a hydrogen

(H) gas jet, as shown in Fig. 81, and hold a piece of spongy platinum (Pt) over it, as shown in Fig. 88, and it will soon get red-hot and then, in turn, it will light the gas. Spongy platinum (Pt) is a powdered form of platinum (Pt), and you can buy it ready to use. It is made by dissolving platinum (Pt) in *aqua regia*, which is a mixture of 1 part of nitric acid (HNO_3) and 3 parts of hydrochloric acid (HCl). Crystals of chloroplatinic acid (H_2PtCl_6) are thus formed; ammonium chloride (NH_4Cl) is then added, which precipitates the platinum (Pt) as ammonium chloroplatinate ($(NH_4)_2PtCl_6$), and on heating the compound it leaves the platinum (Pt) in a powdered form, and this is called *spongy platinum* (Pt).

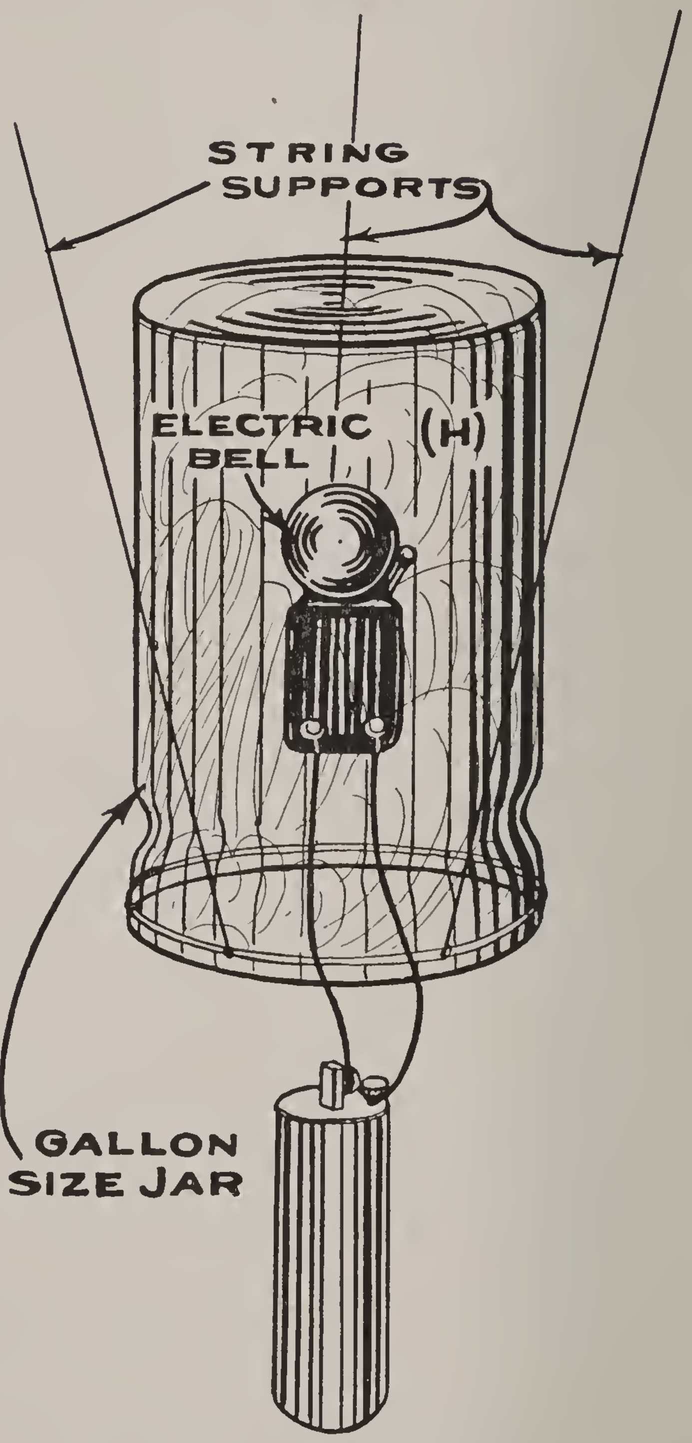

Fig. 89.—An Electric Bell in Hydrogen.

How Hydrogen Acts on Silver Nitrate. Dissolve as much silver nitrate ($AgNO_3$) as you can in a teaspoonful of pure water (H_2O), and with a toothpick write or draw upon a piece of silk; this done, moisten the latter with a damp sponge and then place the silk in a beaker of hydrogen (H), or you can direct a stream of the gas on it with the bladder or rubber-bag apparatus. The hydrogen (H) removes the oxygen (O) of the silver nitrate ($AgNO_3$) and leaves the pure metal on the silk.

Fig. 90.—A Squeaking Head in Hydrogen.

How Hydrogen Acts on Sound. For the following experiments take a gallon glass fruit jar, or a battery jar, fill it with hydrogen (H), and suspend it by means of strings; now connect a dry cell with an electric bell and hold the latter in the jar, as shown in Fig. 89, and it will give out quite a different sound from that which it does when it rings in air.

If you can get one of those squeaking toys that are sold at Christmas time, made in the form of a head, a duck, or a dog, and work it in a jar of hydrogen (H), as shown in Fig. 90, it will give forth a most ridiculous sound which is very funny. In fact any object which will make a sound in air and that can be worked in a jar of hydrogen (*H*) will set up a weird and curious noise.

While this experiment is not an easy one to make, still if you can do it you will cause no end of astonishment. Invert a tin, or copper, wash-boiler and suspend it by its handles. This done, fill it with hydrogen (*H*) and then play an accordion in it. After hearing the wonderful music it makes you will not need to stretch your imagination to conceive what a whole orchestra shut up in a room of hydrogen (*H*) would sound like.

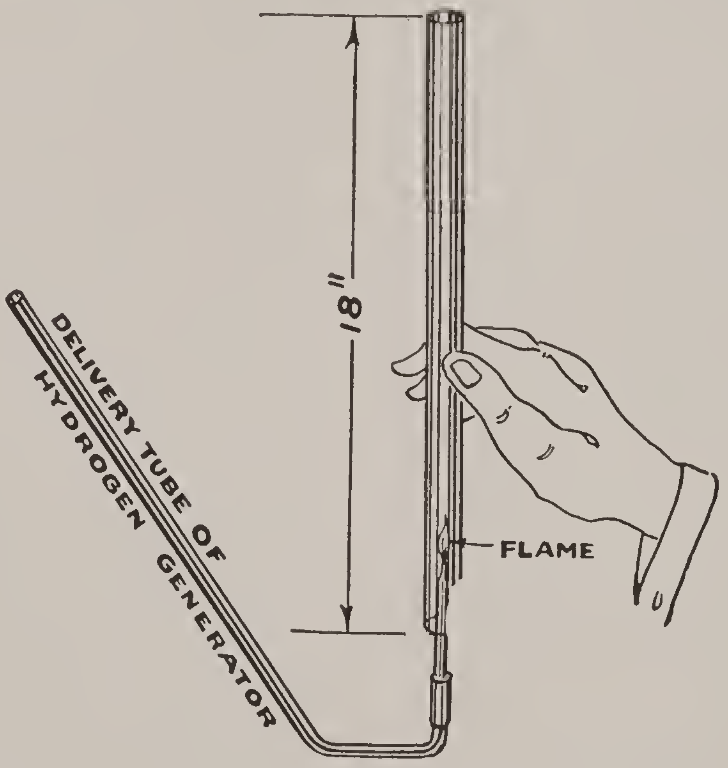

Fig. 91.—A Hydrogen Flame Organ Pipe.

Here is an experiment that has been made but which you are not advised to try. When perfectly pure hydrogen (*H*) is inhaled it is like nitrogen (*N*) in that it kills, not because it is a poison but because it will not support life. A most curious effect is produced on the voice by inhaling some of the gas and then speaking, or singing, while it is being exhaled, for it changes the deep voice of a man into a nasal, piping tone that is both curious and funny. In this experiment, two precautions must be taken, and these are to use absolutely pure hydrogen (*H*), and to keep entirely away from a flame of any kind while inhaling the gas.

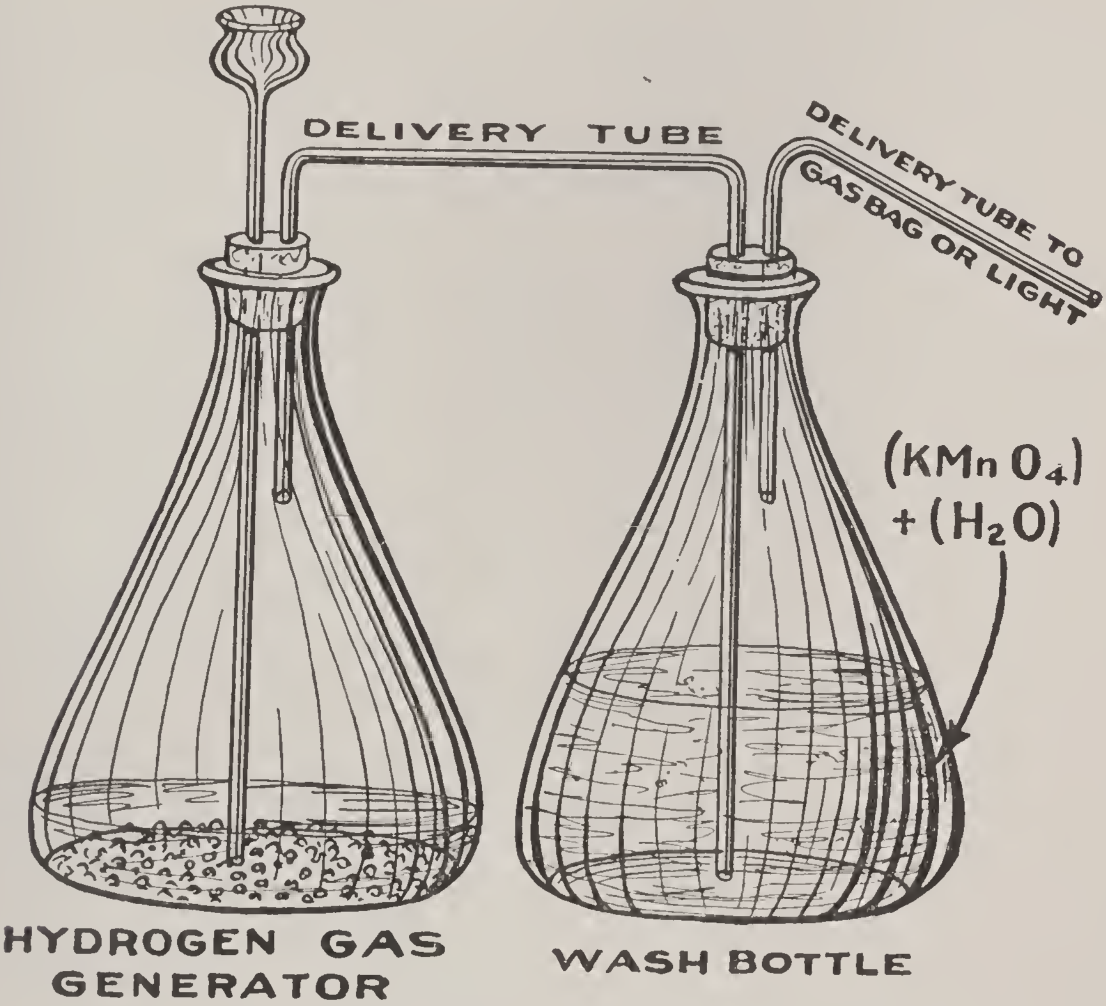

FIG. 92.—How to Purify Hydrogen and Other Gases.

How to Make a Hydrogen-Flame Organ Pipe. For this experiment use the apparatus described under the caption of "How to Make a Hydrogen Flame" and hold a glass tube that has a bore of 1 or 1½ inches and a length of 18 or 20 inches and open at both ends, over a very small flame, as pictured in Fig. 91. Now raise and lower the tube a little at a time, and you will strike a point where it will give out a clear musical note, and then by moving it up or down, different tones will be produced.

How the Experiment Works. When 32 or more vibrations take place in a second, a musical sound is set up that the ear can hear. Now when the flame is made to burn in the tube there will be a large number of regular explosions of the hydrogen (*H*), and this sets up waves in the air, which in turn produce a musical sound.

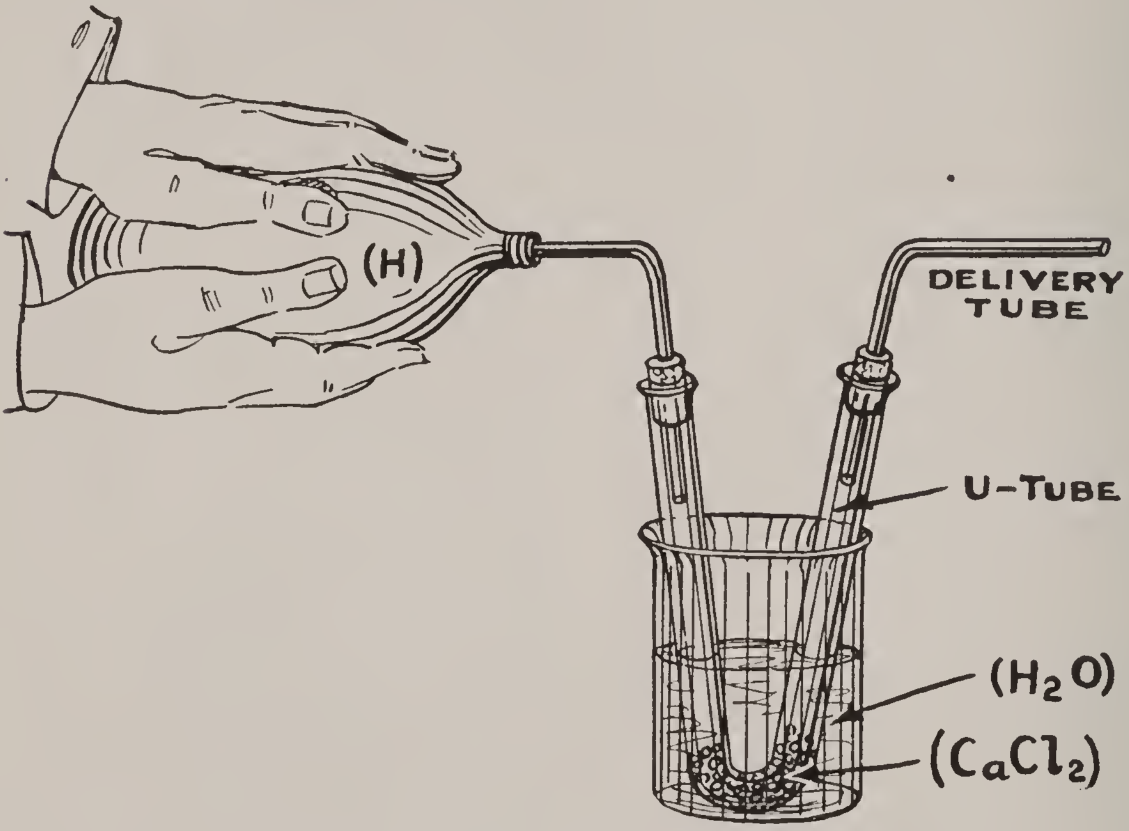

FIG. 93.—How to Dry Hydrogen and Other Gases.

How to Purify Hydrogen Gas. In making experiments with hydrogen (*H*), it is often necessary to have it free from all other substances. To purify it you need only to pass it through a solution made by dissolving 1 ounce of potassium permanganate ($KMnO_4$) in 4 or 5 ounces of water (H_2O). This solution is put in a wash-bottle, that is, a flask or any wide-mouth bottle with a tight-fitting cork in which there are two tubes; the longer one, which reaches below the surface of the solution, is connected with the delivery tube of the generating apparatus, and the purified hydrogen (H) is given off through the short tube, as shown in Fig. 92.

How to Dry Hydrogen. Likewise it is often desirable that the hydrogen (H) after it is purified should be perfectly dry. You can easily extract the moisture from it by first passing it from the wash-bottle connected with the generating apparatus into a bladder, or rubber bag; now connect the latter with one of the ends of a U-tube in which you have placed 1 ounce, or more, of calcium chloride ($CaCl_2$) and the other end with a delivery tube, as shown in Fig. 93. Calcium chloride ($CaCl_2$) is a substance that is very *deliquescent*, that is, it has the power of absorbing large quantities of water (H_2O) and, hence, it extracts whatever moisture there may be in a gas which is either passed through or over it.

CHAPTER VI

A PAIR OF SMELLY GASES

THERE are less than a dozen gases that are known to exist as elements, and the most important of these, namely oxygen (O), nitrogen (N), and hydrogen (H), I have told you about in the chapters that have gone before. There is, however, another gas called chlorine (Cl) that is an element, and as it is widely used I want you to know about it, too.

Now while chlorine (Cl) is not found free in nature, there is an enormous amount of it locked up in various compounds, as for instance in sodium chloride ($NaCl$), that is, *common salt*. As 3/5 by volume of sodium chloride ($NaCl$) is formed of chlorine (Cl) and there is enough salt in the oceans to make a range of mountains as large as the Alps, it will be seen that it is extremely plentiful and, moreover, it is easy to obtain. In turn, chlorine (Cl) forms a large number of useful compounds when mixed or combined with other elements.

Another interesting gas, though very strong-smelling, which I shall tell you of in this chapter, is ammonia (NH_3), but this is a compound and it is made up, as its formula shows of 1 atom of nitrogen (N) combined with 3 atoms of hydrogen (H). Now free hydrogen (H) is very scarce, and while water (H_2O) is made up of 2/3 by volume of this gas, it is not altogether easy to separate it from the oxygen (O) in commercial amounts. But hydrogen (H) is found in large quantities in all living and dead plant and animal matter and also in natural gas and petroleum, and from these compounds it can be easily obtained; hence, ammonia (NH_3) can be cheaply made in large quantities.

Experiment with Chlorine. Chlorine (Cl) is a transparent gas of a yellowish-green color and Davy, who proved that it is an element, gave it this name from the Greek word *chloros*, which means green. Chlorine (Cl) was discovered by Scheele in 1774, and it was he who first made it. This he did by treating black oxide of manganese (MnO_2) with hydrochloric acid (HCl), or *spirit of salt* as it was called in the early days of chemistry.

Chlorine (Cl) is a gas that is nearly 3 times as heavy as air, and it has a pungent, suffocating odor, which is very penetrating. In making experiments with it, be very careful not to inhale it, for it has an irritating effect on the throat, and pure chlorine (Cl) will kill. While chlorine (Cl) will not itself burn, it will support the combustion of some substances. Scheele not only discovered this gas, but he found out most of its characteristics, including the all-important one that when it is mixed with water (H_2O) it will bleach out all kinds of vegetable colors.

The process for making chlorine (Cl) is not nearly so simple as that for oxygen (O), but you can make it for experimental purposes without any trouble. There are three ways by which it can be made, and these are to pass a current of electricity through a solution of sodium chloride ($NaCl$), that is, common salt, and water (H_2O), just as you did the water (H_2O) alone, to separate it into its component gases, as described in the previous chapter; to use hydrogen chloride (HCl), or some other cheap compound that contains the gas, and make the hydrogen (H) in it combine with the oxygen (O) of the air, setting free the chlorine (Cl); and, by the easier process which I shall now describe.

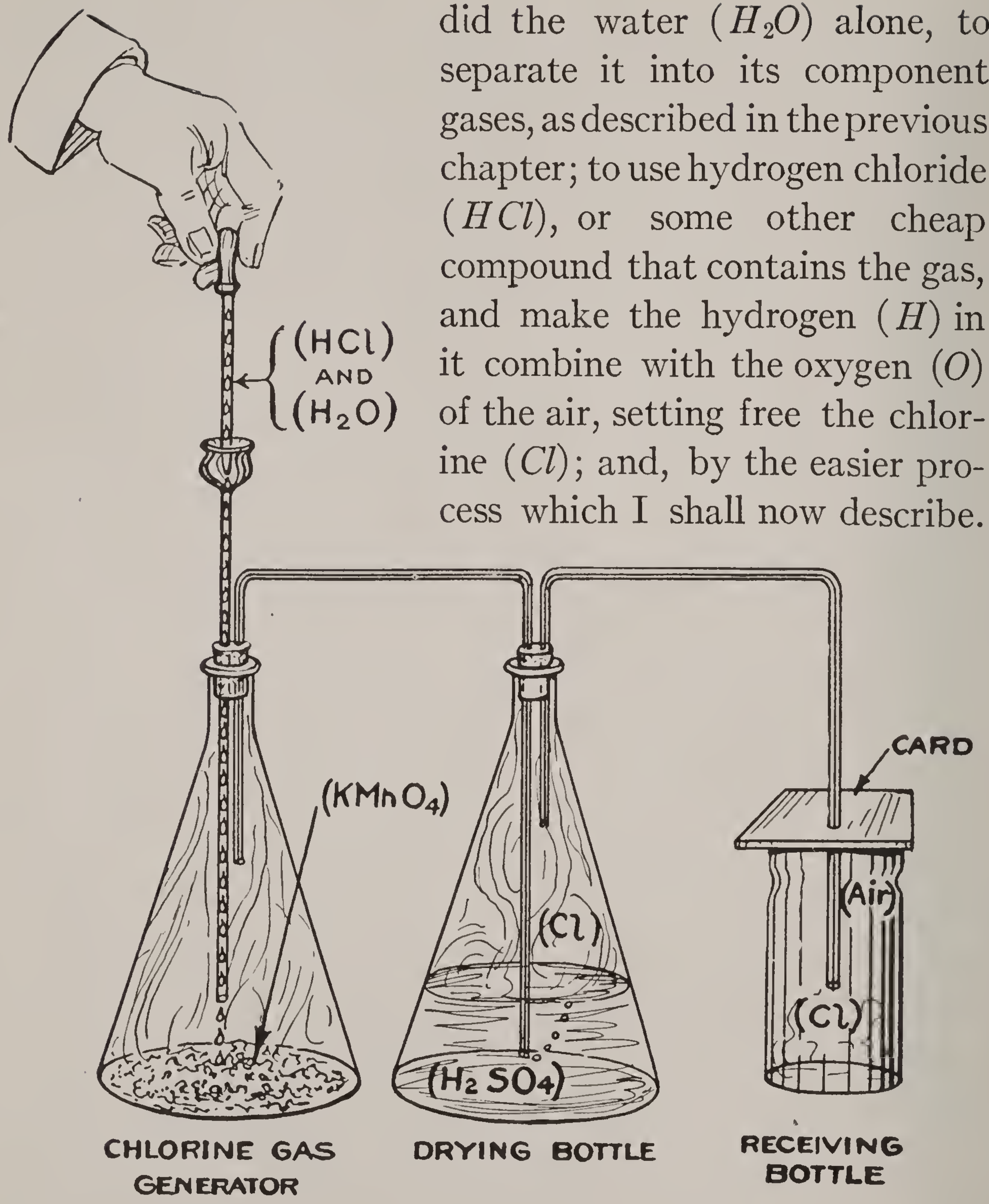

FIG. 94.—Apparatus for Generating Chlorine Gas.

How to Make Chlorine. To make enough chlorine (Cl) to experiment with, all you need is a flask fitted with a funnel and a short delivery tube; connect the latter with the long tube of a wash-bottle while the short tube of the latter leads into a bottle that is to hold the gas, and which stands right side up, as shown in Fig. 94.

Now while chlorine (Cl) is heavier than air, it is a good plan to slip a card over the delivery tube of the wash-bottle and lay it over the mouth of the bottle that is to hold the gas, to prevent its dispersion. If you want the gas to be perfectly dry as well as pure, then you will have to use the wash-bottle as a drying bottle, by filling it about one-third full of concentrated sulphuric acid (H_2SO_4).

Having the apparatus set up, put 1 ounce of potassium permanganate ($KMnO_4$) in the flask and then fit the cork in tight. Now fill a pipette, or medicine dropper, with a solution made of 1 part of water (H_2O) and 3 parts of concentrated hydrochloric acid (HCl), which is the only compound that hydrogen (H) and chlorine (Cl) will form with each other, and let it fall drop by drop into the funnel. A better appliance than the pipette is a funnel with a stopcock in it. The acid is used up very fast and the gas is set free in a quantity large enough to make a constant stream of it flow from the delivery tube, while the bottle is filled by the gas falling to the bottom and forcing the air out at the top.

How the Experiment Works. When the potassium permanganate ($KMnO_4$) comes into contact with the hydrochloric acid (HCl) it forms four substances, and these are water (H_2O), potassium chloride (KCl), manganese chloride ($MnCl_2$), and chlorine (Cl). The water (H_2O) and the chlorides are left behind in the flask, and since the chlorine (Cl) is a gas it passes out of the delivery tube.

NOTE.— In making chlorine (Cl) and experimenting with it, you should do so either in a shed or outdoors, for the gas is not only very bad to breathe but it will destroy the finish on furniture, take the color out of draperies, and spoil the polish on metal work.

How to Test for Chlorine. Pour a little of the gas into a test tube of pure water (H_2O), and then dissolve a crystal of silver nitrate ($AgNO_3$) in a teaspoonful of pure water (H_2O); this done, add a few drops of it to the solution in the test tube. If the solution is chlorine (Cl) in water (H_2O) — which makes hypochlorous acid ($HClO$) — a precipitate will be formed like the curd of milk, and this is silver chloride ($AgCl$); spread this salt out on paper and let the sunlight shine on it and it will turn black.

How Chlorine Acts on Flame. Twist a wire around a match, then light it and lower it into a test tube of chlorine (Cl); the latter will ignite and burn with a dull red flame, then dense fumes will be given off and the flame will soon go out.

Spontaneous Combustion. Cut a strip of filter paper about ½ inch wide and 4 inches long and fold it over lengthwise; dip one end of it into some turpentine that has been warmed, and lower it into a test tube of chlorine (Cl). It will instantly ignite and give off a lot of black smoke.

Moisten a strip of filter paper with concentrated liquid ammonia (NH_3) and lower it into the test tube of chlorine (Cl) and it will ignite spontaneously.

How to Make a Smoke Screen. Here is a way to make smoke that is blacker than the blackest smoke of smoky Pittsburgh. Put 1/3 ounce of wood-alcohol (CH_3OH)

into a long test tube, then add 1 ounce of sulphuric acid (H_2SO_4) and heat them over the flame of your alcohol lamp. This done, grease the mouth of the test tube so that an air-tight joint can be made, and pour some chlorine (Cl) gas into the tube, put a piece of glass on top of it, and shake it well. The next step is to take off the glass and light the gas in the tube, upon which a great cloud of dense black smoke will be formed by the carbon (C) which is set free, as shown in Fig. 95. While the gas is burning, you will hear an ominous noise like that of a miniature earthquake as the flame moves down from the mouth to the bottom of the tube.

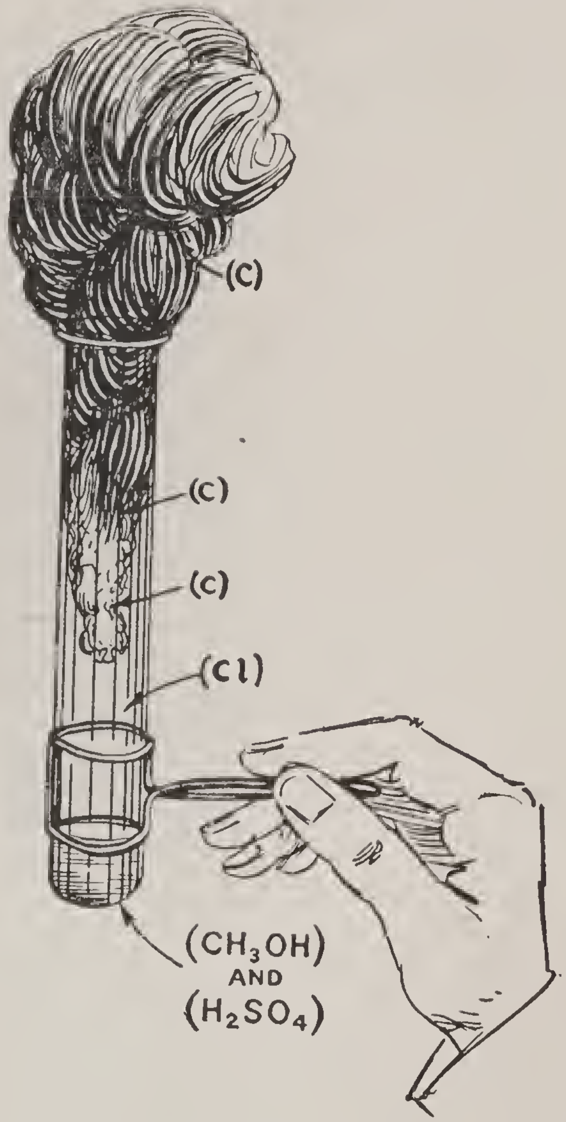

FIG. 95.—How to Make a Smoke Screen.

The Art of Bleaching. Chlorine (Cl) is an active element, and in this respect is very much like oxygen (O); further, it has a larger number of chemical properties than this latter gas has. Now while it is commonly said that chlorine (Cl) *bleaches*, when it is perfectly dry it has no bleaching properties whatever, and before it will bleach it must be brought into contact with water (H_2O). This can be done by adding some chlorine (Cl) to water (H_2O), upon which hypochlorous acid ($HClO$) results, and this has bleaching properties; or the colored piece of goods can be moistened with water (H_2O) first and then dipped into a jar of chlorine (Cl); in either case the colors will fade away until the piece of goods is perfectly white.

FIG. 96.—Making Some Dry Chlorine Gas.

How to Test the Bleaching Power of Chlorine. Put 2 or 3 ounces of calcium chloride ($CaCl_2$), that you have broken up into small lumps, into a perfectly dry jar, then cut a hole in a sheet of blotting paper and put it on top of the jar. Pour some chlorine (Cl) from the jar or bottle containing it into the other one, as shown in Fig. 96; the purpose of the blotting paper is to prevent any water (H_2O) that may have gathered in the chlorine bottle from getting into the dry jar when you pour it out, and the purpose of the calcium chloride ($CaCl_2$) is to absorb whatever water (H_2O) there may be in the gas itself.

Now take a toothpick or a quill and write your name boldly with some ordinary ink on a strip of paper, dry it over a stove to expel any moisture that there may be in it, and while it is yet warm put it into the jar of chlorine (Cl) and cork it up, as shown in Fig. 97, and you will find that the gas has practically no effect on it.

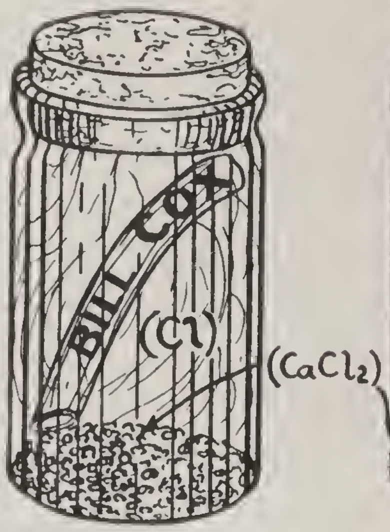

FIG. 97.—The Writing in the Bottle.

FIG. 98.—The Writing Bleached Out.

Having made this experiment, take out the paper, moisten it with water, (H_2O), again put it in the jar of chlorine (Cl) and cork it up, as in Fig. 98, and in a very short time you will find the ink fading, and, finally, it will disappear altogether.

How the Experiment Works. While chlorine (Cl) when it is perfectly dry has no bleaching power, on coming into contact with water (H_2O) it forms hypochlorous acid ($HClO$), as mentioned before. Now when this solution is exposed to the sunlight it decomposes very fast, and hydrogen chloride (HCl), which is a gas, is formed; this remains behind with the water (H_2O), then the solution becomes hydrochloric acid (HCl) (commonly called *muriatic acid*), and the oxygen (O) is set free.

Oxygen (O) is a strong oxidizing agent and, hence, it combines with the organic matter of which the ink is made, and so takes it out of the paper. Oxygen (O) does not act as a bleaching agent under ordinary conditions, but it takes on this power at the instant it is set free from the water (H_2O) by the chlorine (Cl), hence, it is the oxygen (O) which bleaches, and not the chlorine (Cl) itself.

How to Make Red Roses White. Tie the stem of a red rose to a pin pushed through a cork, and then put the latter into the mouth of a bottle of dry chlorine (Cl), as shown in Fig. 99. In the course of a little while the red color will begin to fade away, and finally it will vanish altogether and the rose will be a perfectly white one. The reason that flowers can be bleached in dry chlorine (Cl) is because a very considerable part of them is formed of water (H_2O). Wheat flour is bleached with chlorine (Cl), and this is the way it is made so white.

How to Make Bleaching Powder. Calcium hypochlorite ($CaOCl_2$), or *chloride of lime*, or *bleaching powder*, as it is more often called, is made by passing chlorine (Cl) through calcium hydroxide ($Ca(OH)_2$), or slaked lime, to give it its common name. Chloride of lime ($CaOCl_2$), is an unstable compound which gives up its oxygen (O) freely and leaves calcium chloride ($CaCl_2, 6H_2O$) behind. It is the oxygen (O) that is set free which kills germs in decaying and dead plant and animal matter and it is, therefore, largely used as a *disinfectant* as well as a *bleaching agent.*

To make a little bleaching powder, put a tablespoonful of chloride of lime ($CaOCl_2$) in a beaker and half fill the latter with water (H_2O). Stir it with a glass rod until as much as possible of the lime is dissolved, let it settle, and then pour off the clear solution. Now dip a piece of paper colored with ink, or a piece of muslin with fruit stains on it, into it and they will be bleached out, leaving the paper or muslin perfectly white.

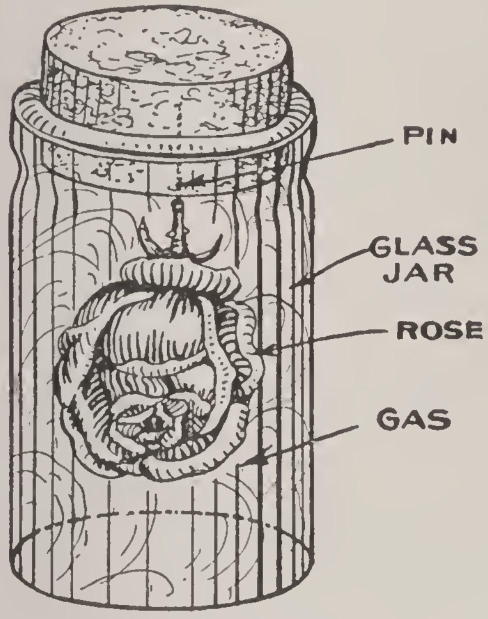

FIG. 99.—To Make a Red Rose White.

How to Make a Bleaching Liquid. Fill a small test tube with water (H_2O), then add a few drops of sulphuric acid (H_2SO_4) and stir with a glass rod; this done, dissolve in the solution as much potassium chlorate ($KClO_3$) as you can put on the point of a knife-blade, and you will find that it has decided bleaching properties.

NOTE.— Do not make this experiment on any larger scale than is given, or you may have an explosion.

How to Make a Bandanna Handkerchief. Fifty years ago a gentleman would be as lost without a *bandanna handkerchief* as he is to-day in Piccadilly without an eye-glass.

But I wonder if you know just what a bandanna handkerchief is. For fear you may not, let me say that it is a very large red silk handkerchief with white spots on it.

When they first became popular, they were made by the very simple expedient of laying red silk handkerchiefs between thin sheets of lead in which there were a number of holes. When a pile of 40 or 50 handkerchiefs and lead plates were laid in this fashion, a solution of chlorine (Cl) and water (H_2O), that is, hypochlorous acid ($HClO$), was poured in the holes in the top plate, and as it seeped through the successive pieces of red silk it took the color out of them and left spots that were perfectly white.

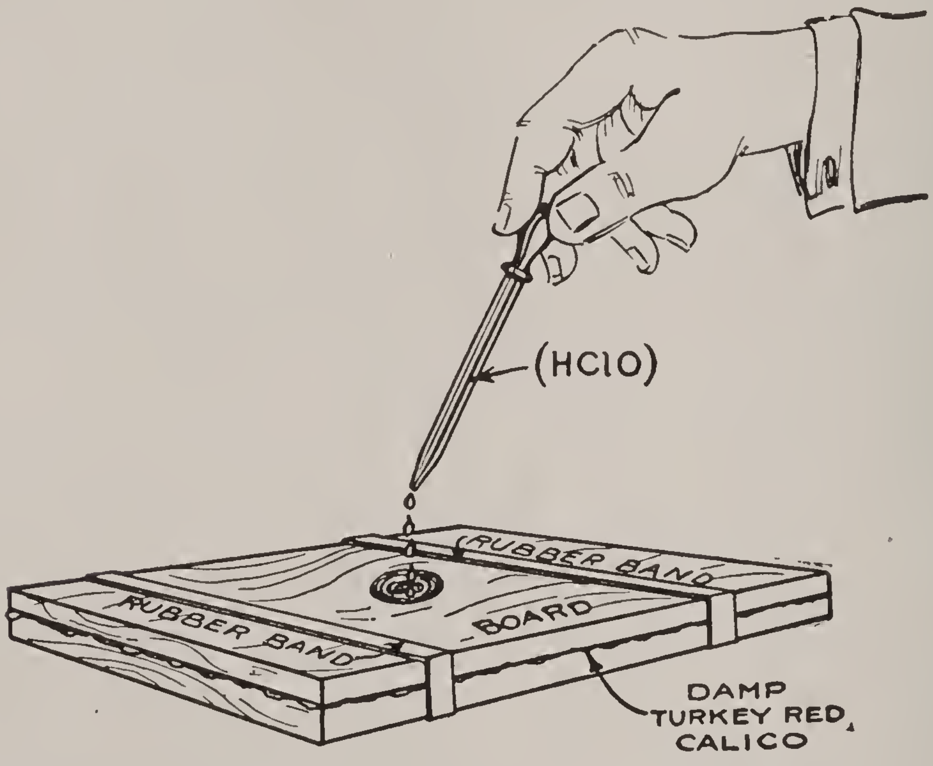

FIG. 100.—How to Make a Bandanna Handkerchief.

You can imitate this process by cutting out a hole an inch in diameter in two strips of cigar-box wood, then placing a piece of calico dyed with Turkey red,[1] which you have moistened, between them, and slipping a strong rubber band over each end, as shown in Fig. 100. This done, drop some hypochlorous acid ($HClO$), which is a bleaching liquid, into the hole of the top board and let it filter through the cloth. Finally, take the calico from the boards and wash it, and you will find a perfectly white spot on it where the liquid came in contact with the colored fibres.

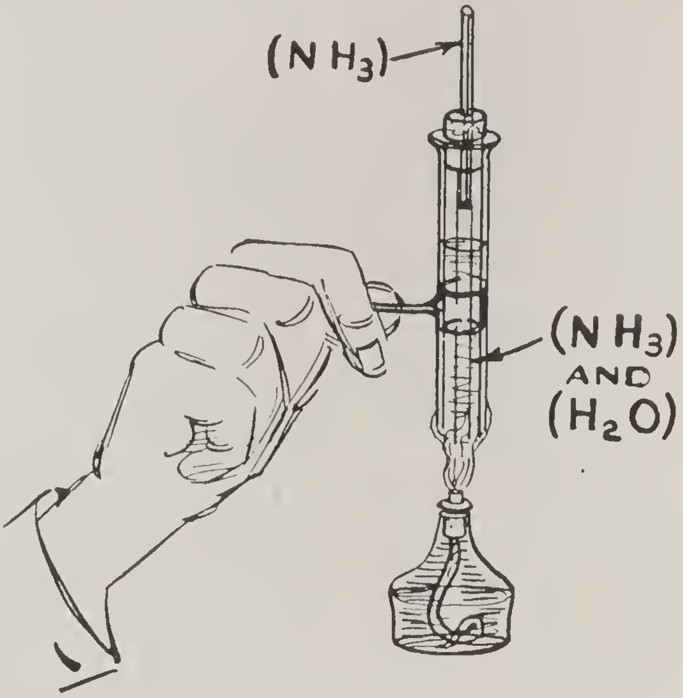

FIG. 101.—Making a Little Ammonia Gas.

Experiment with Ammonia. While the ammonia (NH_4OH) we know so well and use so much of is a liquid formed of ammonia (NH_3), which is a gas, dissolved in water (H_2O), real ammonia (NH_3) is a transparent colorless gas that has a very penetrating, choking odor, and

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[1] Turkey red is a dye produced when alizarin ($C_{14}H_8O_4$) is used with a *mordant* of aluminum sulphate ($Al_2(SO_4)_3.H_2O$).

when inhaled it produces suffocation. Ammonia (NH_3) is only about half as heavy as air, and while it will not burn in the latter, it will burn in oxygen (O).

One of the characteristics of ammonia (NH_3) is that a large amount of it will dissolve in a very small amount of water (H_2O), which is to say that 600 volumes of ammonia (NH_3) can be dissolved in 1 volume of water (H_2O). Now, in the experiments to follow, be sure to keep ammonia (NH_3) gas separate and distinct in your mind from ammonium hydroxide (NH_4OH), which is ammonia (NH_3) gas dissolved in water (H_2O), and is generally called *ammonia*, or more properly, *aqua ammonia*.

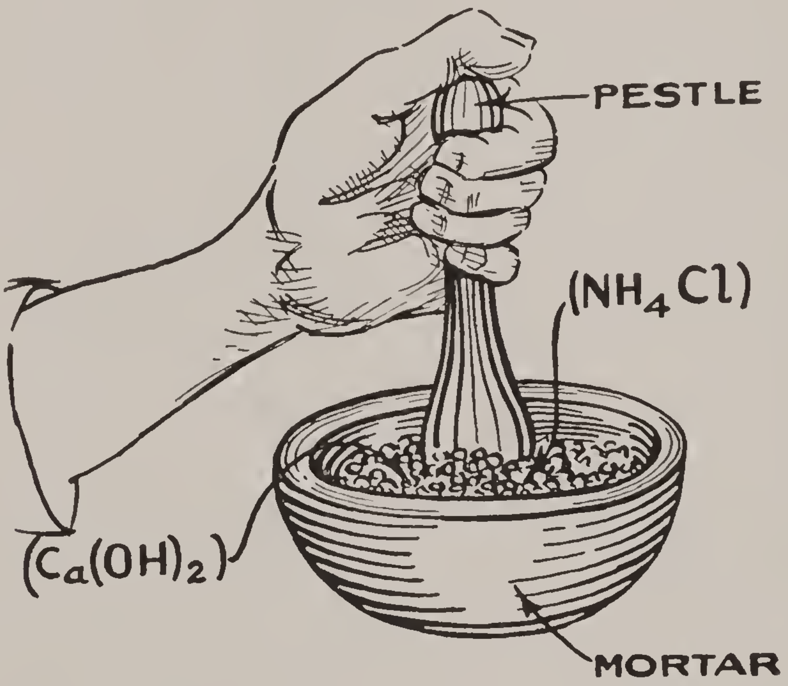

FIG. 102.—Rubbing Up Sal Ammoniac and Slaked Lime in a Mortar.

How to Make a Little Ammonia. Half fill a test tube with *concentrated liquid ammonia*, which is made by dissolving as much ammonia (NH_3) in water (H_2O) as possible, then put a cork into the mouth of it, which is fitted with a delivery tube, as shown in Fig. 101, heat the solution of ammonia (NH_3) over the flame of your alcohol lamp, and ammonia (NH_3) gas will be given off, and this you can detect by cautiously smelling at the opening of the delivery tube.

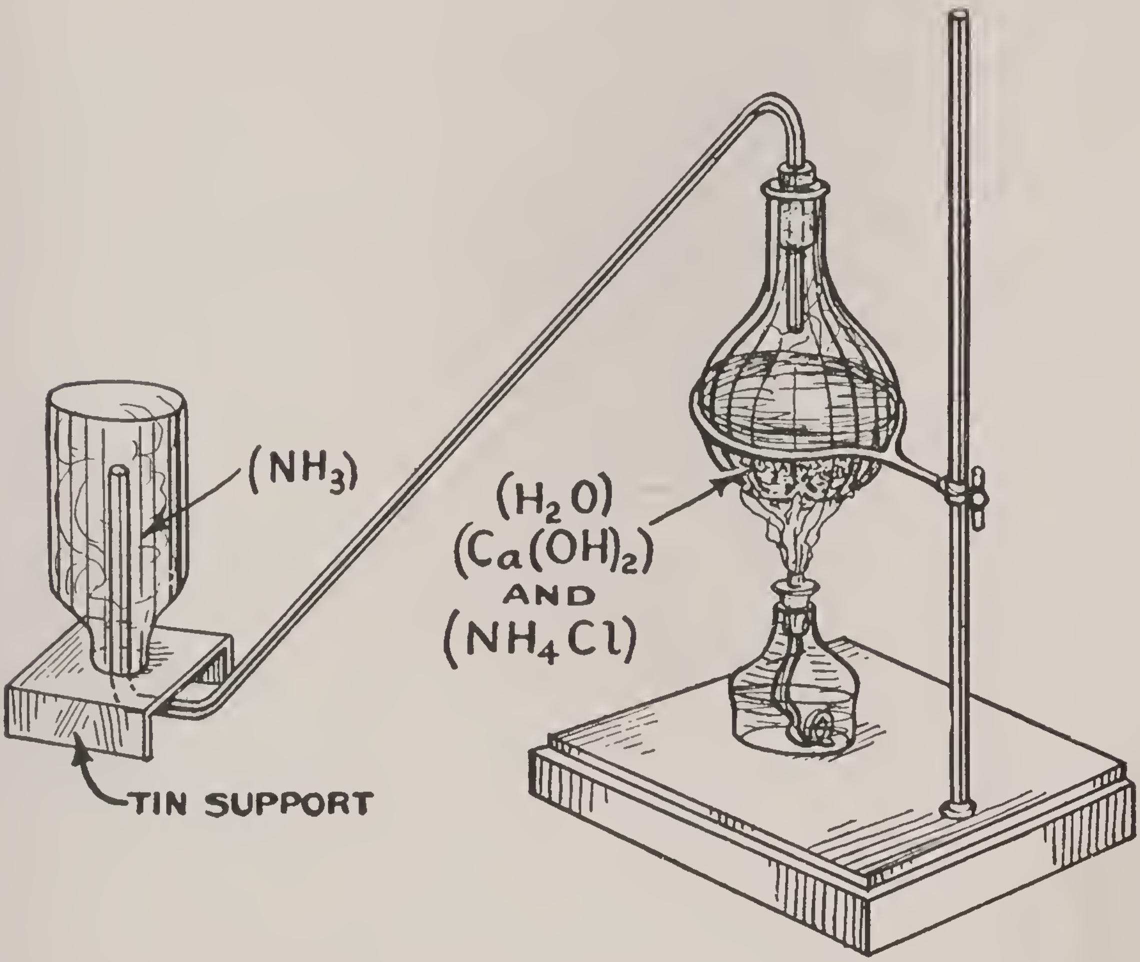

FIG. 103.—Making Ammonia Gas for Experimental Purposes.

How to Make Ammonia on a Larger Scale. Put 1 ounce of ammonium chloride (NH_4Cl), that is, *sal ammoniac*, into a mortar and powder it, as shown in Fig. 102; now powder 1 ounce of calcium hydroxide ($Ca(OH)_2$), which is slaked lime, in a mortar and then mix them well together. This done, put them into a glass flask, add a little warm water (H_2O), and put in a cork that has a delivery tube in it; finally heat it gently over the flame of your alcohol lamp for 10 minutes, as shown in Fig. 103, and ammonia (NH_3) gas will be given off, and, as it is only about half as heavy as air, you can collect it in a large test tube or a bottle by inverting it over the free end of the delivery tube, which should reach nearly to the top of the container.

How the Experiment Works. When the ammonium chloride (NH_4Cl) and calcium hydroxide ($Ca(OH)_2$) are heated together they combine and form calcium chloride ($CaCl_2$), water (H_2O), and ammonia (NH_3) gas, which is given off, or to write it as an equation:

$$\underset{\text{Ammonium chloride}}{NH_4Cl} + \underset{\text{Calcium hydroxide}}{Ca(OH)_2} = \underset{\text{Calcium chloride}}{CaCl_2} + \underset{\text{Water}}{H_2O} + \underset{\text{Ammonia}}{NH_3 \uparrow}$$

To Show How Ammonia Dissolves in Water. Take a strong test tube and fill it with ammonia (NH_3) gas, stand it in a small dish of mercury (Hg) (an individual butter-dish will do) so that the gas will be sealed in the tube air-tight, as shown in Fig. 104. Now set the dish and the tube in a larger glass dish and nearly fill the latter with water (H_2O); these preliminaries attended to, lift the test tube so that its mouth will be just above the surface of the mercury (Hg), but not out of the water (H_2O), and the latter will rush up into the tube and nearly fill it, as shown in Fig. 105.

How the Experiment Works. This curious action is due to the fact that the instant the mouth of the test tube is lifted above the mercury (Hg), the water (H_2O) enters it and absorbs about all of the ammonia (NH_3) gas that fills the tube; this action leaves a vacuum in the tube and the pressure of the outside air on the water (H_2O) in the dish forces the latter up into it.

How to Make an Ammonia-Operated Fountain. For this experiment, which is one in physics as well as in chemistry, use a round flask, and through the cork of it fit a short piece of tube with one end drawn to a nozzle. Now fill the flask with ammonia (NH_3) gas and put the cork into it tight, with the nozzle end of the pipette up into the neck; color some water (H_2O) with a little aniline dye, either red or blue, and then dip the lower and larger end of the pipette into it, as is shown in Fig. 106.

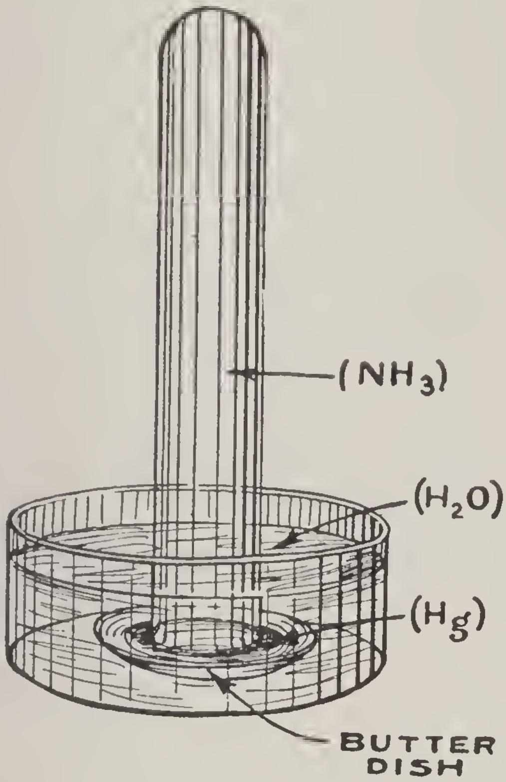

FIG. 104.—The Test Tube Sealed by Mercury.

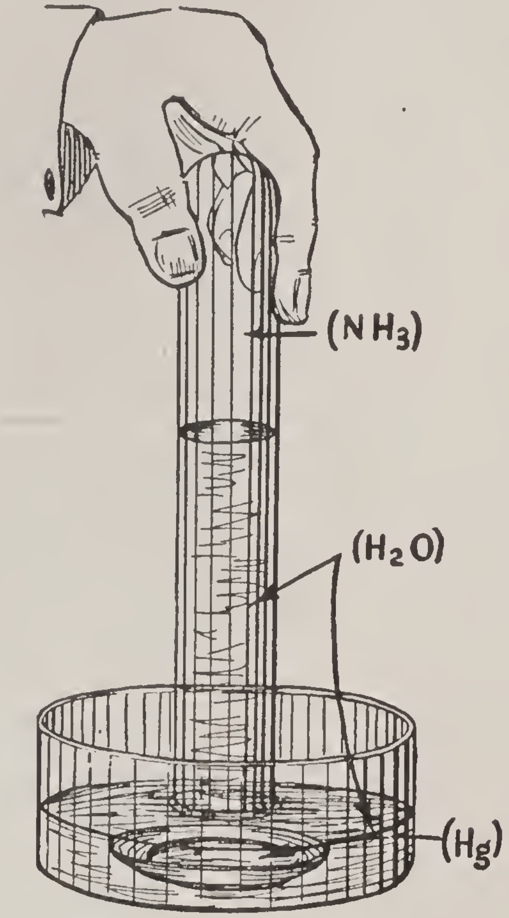

FIG. 105.—The Test Tube Lifted from the Mercury.

Instantly the ammonia (NH_3) gas in the flask will be absorbed by the water (H_2O) and produce a vacuum in the flask. This causes the pressure of the air on the water (H_2O) outside to force it up through the nozzle, after which it will fall in a spray.

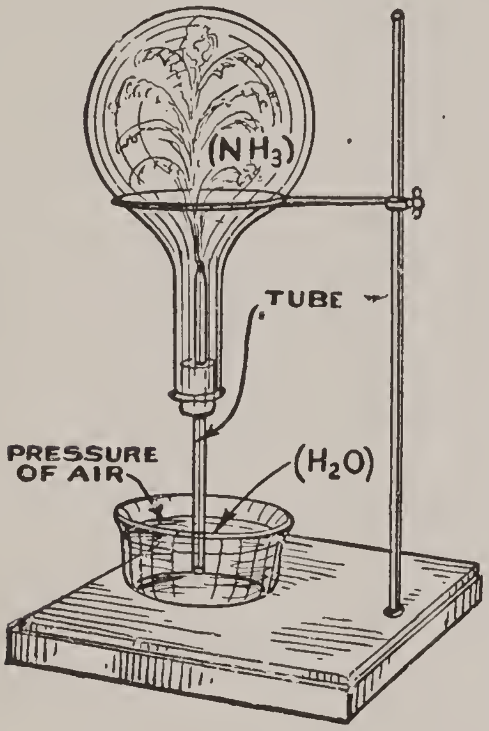

FIG. 106.—An Ammonia-Operated Fountain.

How to Make Concentrated Liquid Ammonia. Liquid ammonia (NH_3) is ammonia (NH_3) gas *liquefied*, which condition is brought about in the same way that air is liquefied, and this is by heat extraction and pressure, but what we ordinarily call liquid ammonia (NH_3) is, as I explained before, simply ammonia (NH_3) gas dissolved in water (H_2O), and *concentrated liquid ammonia* is water (H_2O) in which the largest possible amount of ammonia (NH_3) gas is dissolved.

You can make a small quantity of concentrated liquid ammonia (NH_4OH) by putting a little distilled water (H_2O) into a U-tube and setting this in a beaker of ice-water (H_2O), as shown in Fig. 107. One end of the U-tube is connected with the delivery pipe of your ammonia (NH_3) generating apparatus that is shown in Fig. 103, and the other end of it is closed with a cork in which you have inserted the tube of a pipette. As water (H_2O) increases from 100 volumes in its normal condition to 175 volumes when it is saturated with ammonia (NH_3) gas, the U-tube must not be more than one-third full of water (H_2O) to start with.

How the Experiment Works. When the ammonia (NH_3) is passed *very slowly* into the U-tube the water (H_2O) will rise in the arm of it that contains the pipette, and the pressure of the gas will make the water (H_2O) absorb the largest possible amount, and a saturated solution will result. You will know when the water (H_2O) has absorbed all the ammonia (NH_3) gas it can by the expansion of it, causing it almost to touch the cork that

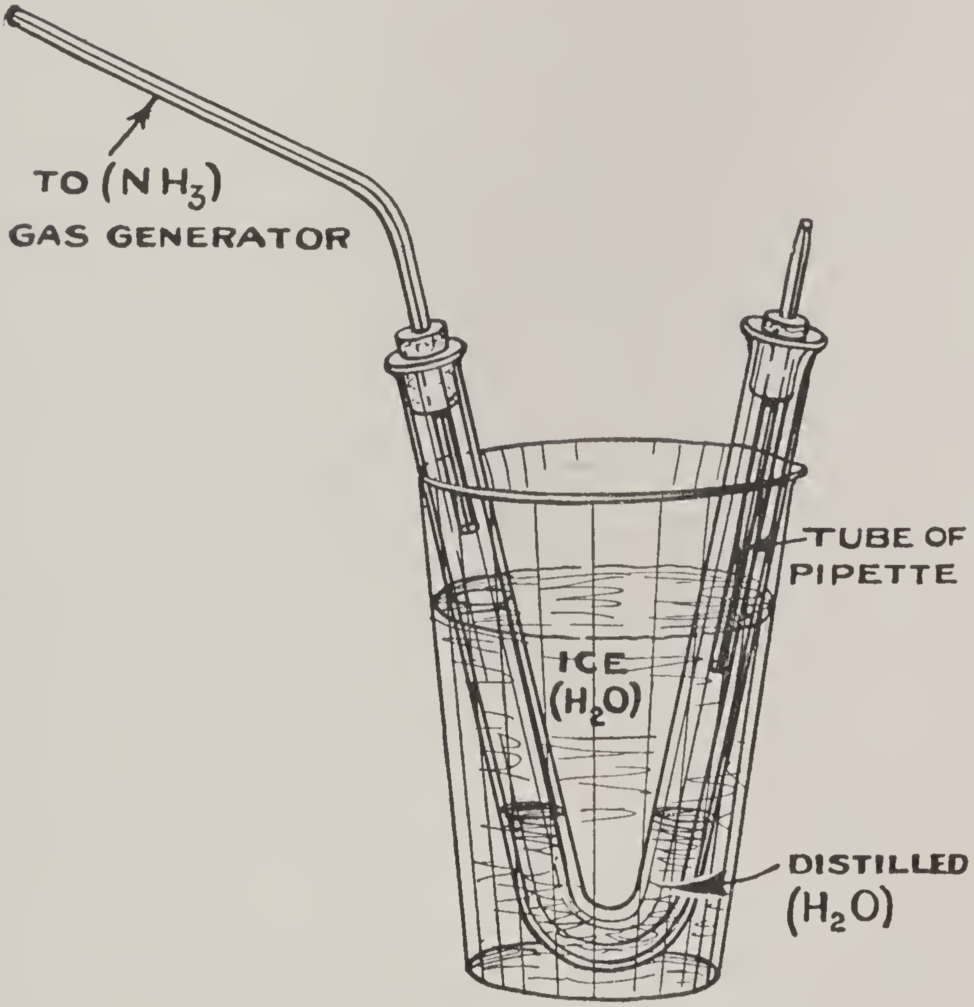

FIG. 107.—Apparatus for Making Concentrated Liquid Ammonia.

has the pipette in it. The resulting solution will be concentrated liquid ammonia (NH_4OH), and it will be very strong indeed.

An Experiment with Concentrated Liquid Ammonia. Put some concentrated liquid ammonia (NH_4OH) into a test tube and set this in a beaker of melting ice (H_2O), and it will be cooled to a temperature of 32 degrees Fahrenheit, which is the freezing point. When it is thoroughly chilled, pour it into a beaker and set this in a warm room; now when it reaches a temperature of 62 degrees Fahrenheit it will begin to boil and give off ammonia (NH_3) gas; as soon as it ceases boiling and giving off gas, grip the beaker with the palm of your hand, as shown in Fig. 108, and the heat of the hand will make the liquid ammonia (NH_4OH) boil again and give off still more gas.

FIG. 108.—Boiling Ammonia with the Heat of Your Hand.

Some Uses of Aqua Ammonia. Ammonia (NH_4OH) is used in small amounts in the household as a remedy for headache (aspirin or phenacetine is better), for polishing metals, for cleaning clothes, and for softening water. It is used in larger amounts commercially for refrigeration, in which the gas is liquefied by pressure and cooled by running it through a coiled pipe in cold water.

The liquid ammonia (NH_4OH) then drips through a nozzle in the end of the pipe, causing evaporation. This takes the heat out of a salt solution, or brine, in which there are suspended cans that are filled with pure water; in this way the temperature of the brine is reduced below the freezing point and the water (H_2O) in the cans freezes into cakes of ice (H_2O). Finally, ammonia is used for making compounds of various kinds, including fertilizers.

CHAPTER VII.

ACIDS, THE GREAT SOLVENTS

WHILE water (H_2O) is the greatest of all solvents, still there are many substances which it will not dissolve, and acids of different kinds must be used instead. All acids are alike in at least four respects, and these are that they have a sour taste, change blue litmus paper red, contain hydrogen (H), and, finally, metals dissolve in them, causing the acids to give up their hydrogen (H).

Acids are usually formed of gases which are dissolved in water (H_2O), and acid solutions of this kind will dissolve metals. Weak acids which are called dilute acids, contain as high as 80 per cent of water (H_2O); stronger acids, called commercial acids, have about 7 per cent of water (H_2O) in them, while concentrated acids have the smallest possible amount of water (H_2O) in them. It is very seldom that the water (H_2O) which forms a part of an acid has any effect on the substances to be acted upon and, hence, it is not taken into consideration in the reactions.

The most useful acid is sulphuric acid (H_2SO_4), then comes hydrochloric acid (HCl), and this is followed by nitric acid (HNO_3); and then there are more than 50 other kinds which have a less extensive use. Now, like all other chemicals, you can buy the acids you need a great deal cheaper than you can make them yourself, but still you should make them for the experience that will accrue to you.

About Sulphuric Acid. Sulphuric acid (H_2SO_4), the common name of which is *oil of vitriol*, is a thick, oily liquid that is nearly twice as heavy as water (H_2O). It is the most difficult of all the common acids to make in the laboratory, but with a little patience you can do it. The start has to be made with sulphur dioxide (SO_2), a colorless gas that is more than twice as heavy as air, and 80 volumes of it will dissolve in 1 volume of water (H_2O). This gas is easily made when sulphur (S) is heated in the air, causing it to combine with the oxygen (O) of the latter.

Next, the sulphur dioxide (SO_2) must be converted into sulphur trioxide (SO_3), which is a colorless, volatile liquid, and this is done by heating the sulphur dioxide (SO_2) and more oxygen (O) together at a high temperature. Another atom of oxygen (O) then combines with each molecule of it and so converts it into a different substance. The vapor of the sulphur trioxide (SO_3) is then conducted to a vessel which is kept cold, and it will liquefy into concentrated sulphuric acid (H_2SO_4).

The Easiest Way to Make Sulphuric Acid. Take a dozen pieces of cotton thread, each about 3 inches long, dip them into melted sulphur (S), and when they are cold tie them to one end of an iron wire about 6 inches long. Now push the other end of the wire into a cork and then put as much water (H_2O) as will fill a small test tube into a pint bottle. This done, light the sulphur (S) on the end of the wire, push the wire into the bottle, and cork it up, as shown in Fig. 109.

When the vapor of the burning sulphur (S) fills the bottle, dip a thin pine stick into some strong nitric acid (HNO_3) and hold it in it, as shown in Fig. 110. Instantly the nitric acid (HNO_3) will decompose and the nitric oxide (N_2O) which is formed will combine with the oxygen (O) and form nitrogen peroxide (NO_2). Let the bottle stand for 10 to 15 minutes and then shake it. The vapors will be absorbed by the water (H_2O). The solution that results is dilute sulphuric acid (H_2SO_4), and you can test it by

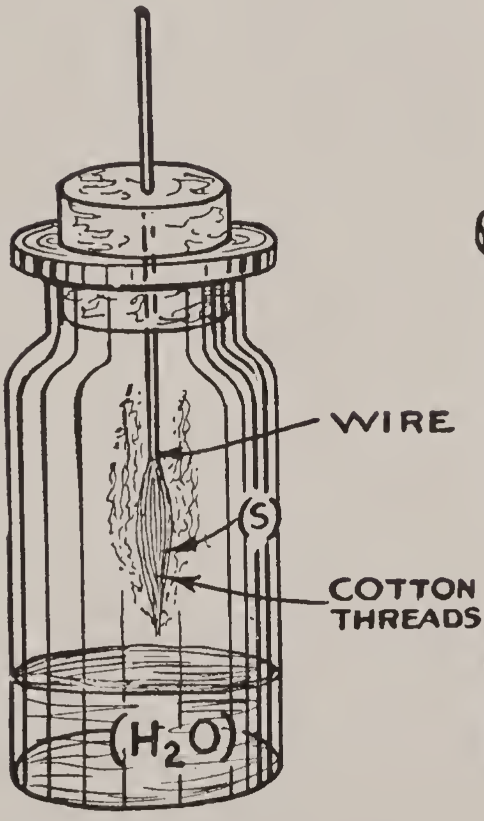

FIG. 109.—Sulphur Burning in a Bottle.

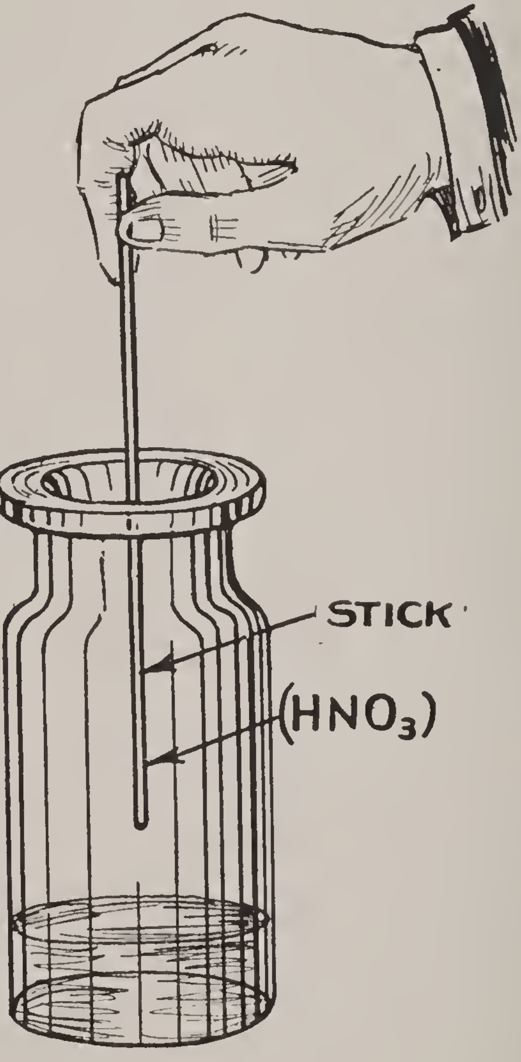

FIG. 110.—Introducing the Nitric Acid.

dipping a piece of blue litmus paper in it. The paper will then turn red.

A Better Way to Make Sulphuric Acid. Here is another way to make a little dilute sulphuric acid (H_2SO_4), and a better one than that just described. Put 1 ounce of potassium nitrate (KNO_3), or *nitre*, as it is called, and 2 ounces of sulphur (S) into a little cup and set it on a block of wood, or other support, which is about 1 inch high, in a saucer of water (H_2O). Now light the mixture and set a larger beaker over the cup in the water, and it will be air-tight, as shown in Fig. 111. The action produces nitrogen peroxide (NO_2), as in the foregoing experiment, and this is absorbed by the water (H_2O), thus forming dilute sulphuric acid (H_2SO_4). To prove it is an acid, test it with a piece of litmus paper as before.

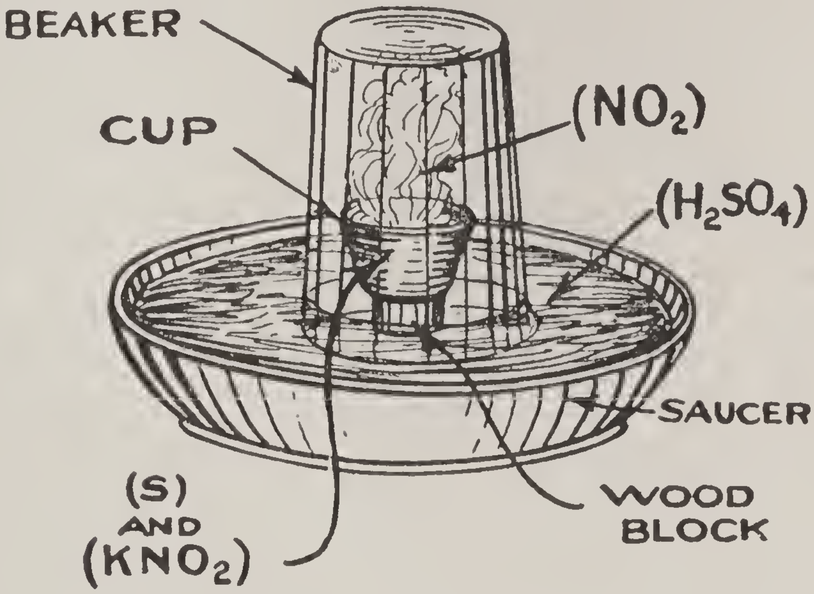

FIG. 111.—A Better Way to Make Dilute Sulphuric Acid.

Another Method for Making Sulphuric Acid. Take 1 ounce of sulphur trioxide (SO_3), which is a compound formed of white silky crystals that look very much like the fibres of asbestos. This must be kept in a bottle with a glass stopper until you are ready to use it, as it fumes strongly when it comes in contact with the air. Put the sulphur trioxide (SO_3) into 1 pint of water (H_2O), and it will dissolve, in doing which it will make a hissing sound and set up a large amount of heat. The resulting solution formed by the reaction is sulphuric acid (H_2SO_4).

A Laboratory Method for Making Sulphuric Acid. To make sulphuric acid (H_2SO_4) by the process that is generally used in school laboratories, you have to begin with oxygen (O) and sulphur dioxide (SO_2), then change these into sulphur trioxide (SO_3), and, finally, dissolve this in water (H_2O). Sulphuric acid (H_2SO_4) results.

How to Make Sulphur Dioxide. Put 1 ounce of copper turnings (Cu) into a pint flask and pour 2 ounces of concentrated sulphuric acid (H_2SO_4) on them; this done, put a cork that has a delivery tube in it in the flask and heat the latter gently, and sulphur dioxide (SO_2) will be given off. Just as soon as the gas begins to pass over, raise the flask above the flame high enough so that the gas will flow from the delivery tube in a steady stream. Couple the delivery tube with a wash-bottle, and then you are ready to connect it with the apparatus for making sulphur trioxide (SO_3). See Fig. 112.

How to Make Sulphur Trioxide. Sulphur trioxide (SO_3), as its formula shows, has 1 more atom of oxygen in each of its molecules than sulphur dioxide (SO_2) has, and to add this extra atom of oxygen (O) you need the following piece of apparatus. Having your oxygen (O)-generat-

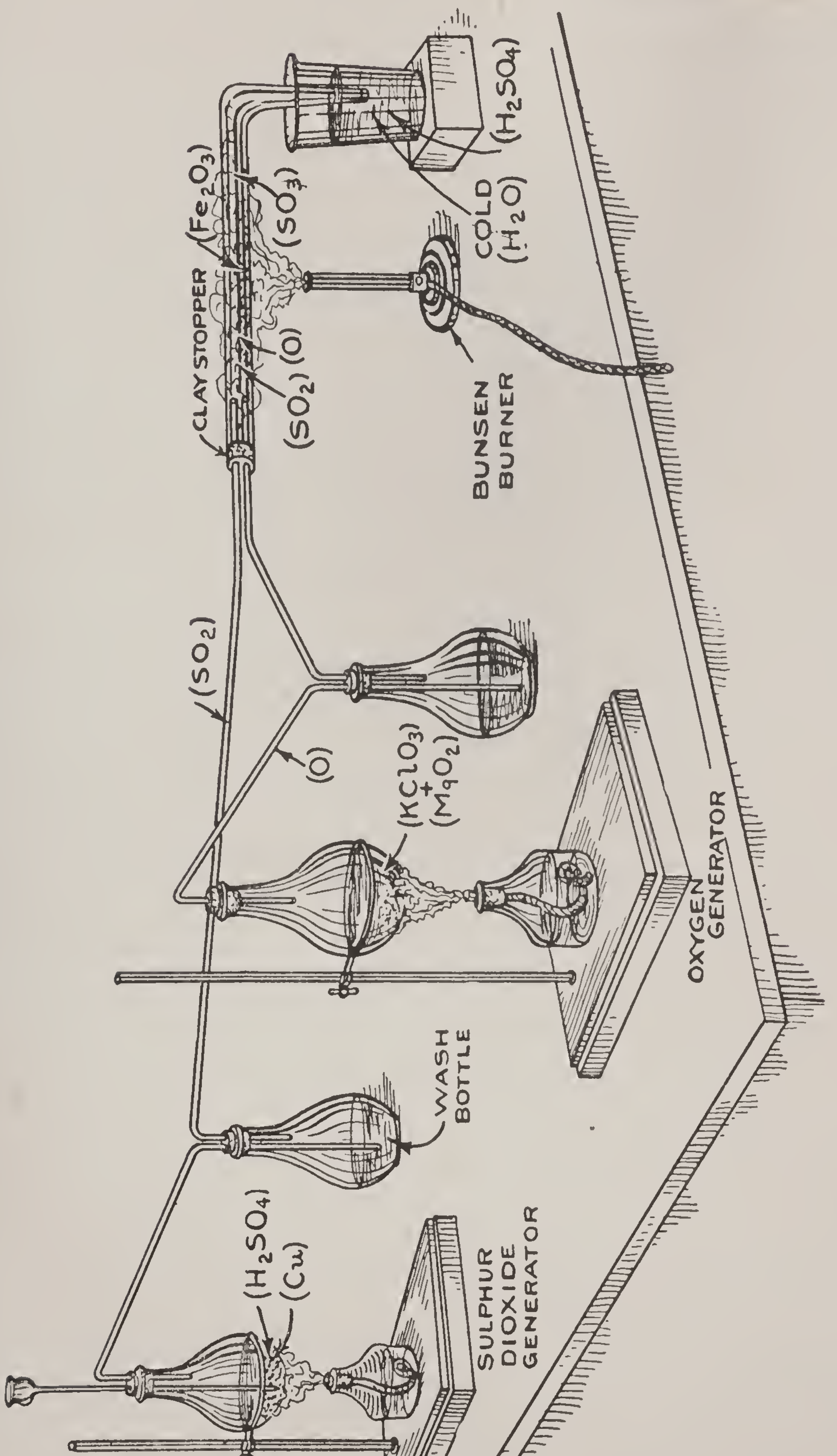

FIG. 112.—Laboratory Method for Making Sulphuric Acid.

ing apparatus, which I described in Chapter III, and the sulphur-dioxide (SO_2) apparatus, just described, set up and ready for action, put 2 ounces of ferric oxide[1] (Fe_2O_3), or iron oxide, as it is usually called and which is common iron rust, into a glass or iron tube ½ inch or ¾ inch in diameter and 8 inches long, one end of which is drawn out to a point to form a nozzle.

Now push the delivery tubes of both the gas generators into the mouth of the glass or iron tube and close it up with a piece of soft clay; now heat the iron oxide (Fe_2O_3) with a Bunsen burner, or, better, a couple of them (an alcohol flame is not hot enough), and then start the generators going and the oxygen (O) and sulphur dioxide (SO_2) will combine and form sulphur trioxide (SO_3), which will pass off through the nozzle of the delivery tube.

How to Make Sulphuric Acid. To make sulphuric acid (H_2SO_4) you need only to set the nozzle end of the tube into a beaker containing a little ice-water (H_2O), and as the sulphur trioxide (SO_3) flows out through the former it condenses into sulphuric acid (H_2SO_4). The complete apparatus for generating the oxygen (O) and sulphur dioxide (SO_2), changing these into sulphur trioxide (SO_3), and finally collecting this as sulphuric acid (H_2SO_4), is shown in Fig. 112.

How the Experiment Works. There are certain sub-

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[1] This acts as a *contact agent*, or *catalytic agent*, which is explained in Chapter III under the caption of "A Way to Make More Oxygen." In making sulphur trioxide (SO_3), *platinized asbestos* is a better contact agent than ferric oxide ($Fe\,O_3$), and it is also much more expensive.

Platinized asbestos is made by soaking the fibers of *asbestos*, or *mineral wool*, as it is commonly called, in chloroplatinic acid ($H_2\ PtCl_3$); this acid is made by dissolving platinum (Pt) in *aqua regia*, and this, in turn, is made by mixing hydrochloric acid (HCl) with nitric acid (HNO_3).

stances, such as oxygen (O) and sulphur dioxide (SO_2), which combine very much better with each other when they are brought into contact with some other substance as, for instance, iron oxide (Fe_2O_3); the latter, curiously enough, does not in any way combine with them and, hence, it remains itself unchanged. Such substances are called *catalyzers* or *catalytic agents*, while the process itself is called *catalysis*. This way of making sulphuric acid (H_2SO_4) is known as the *contact process*.

The reactions that take place in making sulphuric acid (H_2SO_4) by the contact process are these: The copper (Cu) acting on the sulphuric acid (H_2SO_4) gives copper sulphate ($CuSO_4$), or blue vitriol, as it is called, and water (H_2O) which substances remain behind in the flask while the sulphur dioxide (SO_2) gas passes off through the delivery tube. The reaction may be more easily shown by this equation:

$$\underset{\substack{\text{Sulphuric}\\\text{acid}}}{H_2SO_4} + \underset{\text{Copper}}{Cu} = \underset{\substack{\text{Copper}\\\text{sulphate}}}{CuSO_4} + \underset{\text{Water}}{H_2O} + \underset{\substack{\text{Sulphur}\\\text{dioxide}}}{SO_2} \quad \underset{\text{A gas}}{\uparrow}$$

The arrow pointing upward shows that the resultant produ[illegible] is a gas.

EXPERIMENTS WITH SULPHURIC ACID.

How to Change Sugar into Carbon. Put a couple of pieces of lump sugar ($C_{12}H_{22}O_{11}$) in a beaker and pour a tablespoonful of boiling water (H_2O) on them; now add a few drops of sulphuric acid (H_2SO_4) to the solution and it will begin to boil. The hydrogen (H) and the oxygen

(O) of the sugar ($C_{12}H_{22}O_{11}$) combine and form water (H_2O), while the carbon (C) in it is left behind.

How to Write Indelibly on Cotton Goods. Write your name on a piece of white muslin with dilute sulphuric acid (H_2SO_4) and then quickly wash it out well, and there will be no apparent change in the muslin. This done, heat the muslin so that the water (H_2O) in it is driven off, while the trace of sulphuric acid (H_2SO_4) that still remains in the fibres decomposes them and makes them black, and no amount of washing will ever take the color out.

How to Make Copperas. To make ferrous sulphate ($FeSO_4$) or *copperas*, *green vitriol*, or *iron sulphate*, as it is variously known, put a dozen iron (Fe) shingle nails in a flask and add enough dilute sulphuric acid (H_2SO_4) (1 part of water (H_2O) to 5 parts of acid) to cover them. Now warm the flask a little over the flame of your alcohol lamp, and when all the hydrogen (H) has been set free, pour the clear liquid off into a beaker. This done, add a few drops of dilute sulphuric aicd (H_2SO_4) to it, then heat it until half of it has boiled away and let it cool, and green crystals of copperas ($FeSO_4$) will be formed in it. Finally pour off the liquid and lay the crystals on a sheet of blotting paper to dry.

How to Make Blue Vitriol. Cupric sulphate ($CuSO_4$), or *copper sulphate*, or *blue vitriol*, as it is more often called, comes in the form of large blue crystals. Like the crystals of copperas ($FeSO_4$) above described, blue vitriol ($CuSO_4$) contains a large amount of water of crystallization (see Chapter IV), and when the crystals lose this they lose their color, and become what is called *white vitriol*. But just as soon as water (H_2O) touches them again, they take on a blue color. To make copper sulphate ($CuSO_4$) crystals, let some dilute sulphuric acid (H_2SO_4) trickle over copper (Cu) borings, or, better, granulated copper (Cu), in the presence of air.

How to Make Epsom Salts. The chemical name for *Epsom salts* is *magnesium sulphate* ($MgSO_4$), and to make a little of it dissolve 1 ounce of magnesium carbonate ($MgCO_3$) — that is *magnesite* — in dilute sulphuric acid (H_2SO_4), filter it through filter paper placed in a funnel, and catch it in a small porcelain dish. Now heat this gently over the flame of your lamp or burner until the crystals separate from it, and then dry them on blotting paper. You will then have a dose of the famous Epsom salts that have long been used as a laxative.

About Nitric Acid. When sodium nitrate ($NaNO_3$) or *Chili saltpeter*, as it is commonly called, or potassium nitrate (KNO_3), which is *Bengal saltpeter*, is acted upon by any kind of an acid nitric acid (HNO_3) is formed. Since sodium nitrate ($NaNO_3$) is the cheapest nitrate and sulphuric acid (H_2SO_4) is the cheapest acid, they are always used for making nitric acid (HNO_3).

Pure nitric acid (HNO_3) is a colorless liquid that fumes when it is set in the open air. *Commercial* nitric acid (HNO_3) contains 68 per cent of the acid and the rest is water (H_2O), and when you buy *commercial* nitric acid (HNO_3), this is the strength you get. If you really want concentrated nitric acid (HNO_3), then you must buy *pure* acid.

How to Make Nitric Acid. It is as easy to make nitric acid (HNO_3) as it is hard to make sulphuric acid (H_2SO_4). In fact, if you have sulphuric acid (H_2SO_4) to start with, it is easy to make almost any other kind of acid. To make some nitric acid (HNO_3), put 1 ounce of sodium nitrate ($NaNO_3$), which is a salt of the metal sodium (Na), and

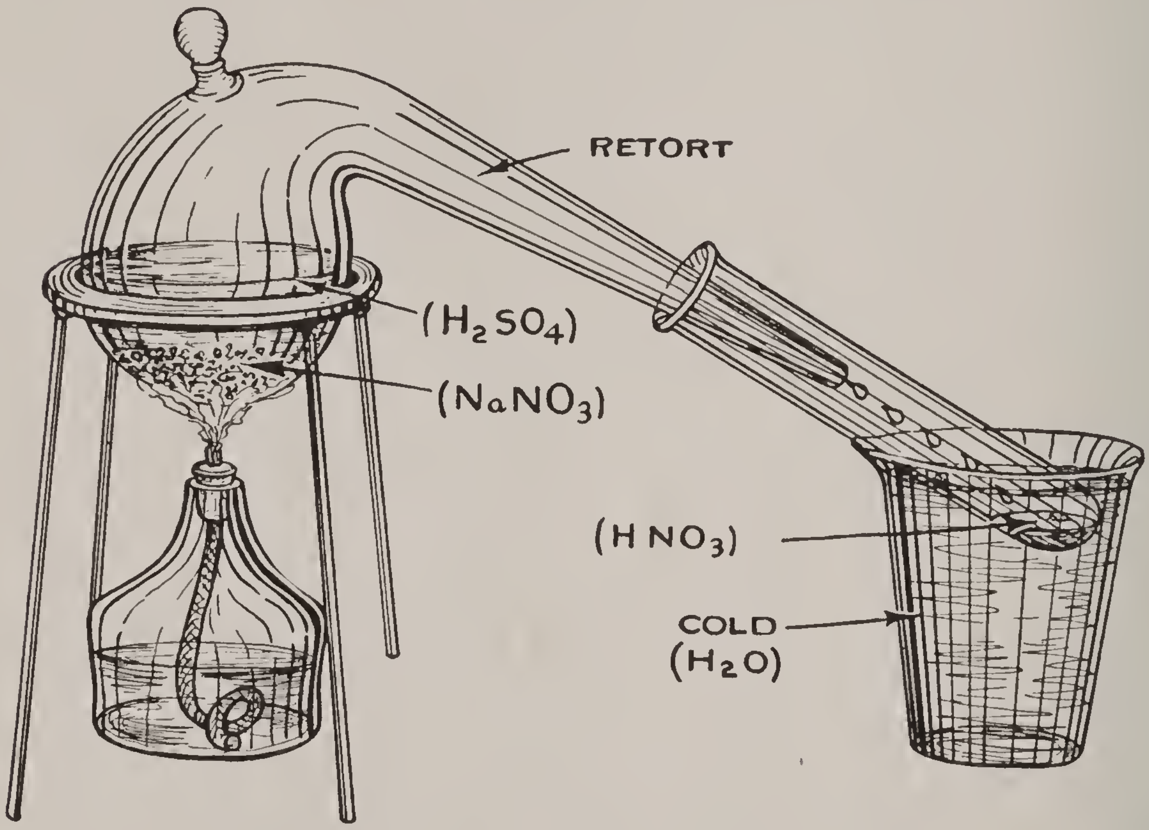

FIG. 113.—How to Make Nitric Acid.

½ ounce of concentrated sulphuric acid (H_2SO_4) into a retort and set your lamp or burner under it; this done, put the mouth of the delivery tube into a test tube and set this in a beaker of cold water, all of which is shown in Fig. 113. Now light the lamp or burner, and the mixture in the retort will give off nitric acid (HNO_3) as a vapor, and this is condensed in the test tube.

How the Experiment Works. After you have made the nitric acid (HNO_3) you will find a substance, or *residue*, as it is called, left over in the retort. Now when the sodium nitrate ($NaNO_3$) and the sulphuric acid (H_2SO_4) are heated together they combine and form sodium sulphate ($NaHSO_4$), which is the residue, and nitric acid (HNO_3), which passes over as a vapor and is condensed into a liquid in the test tube. The reaction can be more easily shown by the equation:

$NaNO_3$	$+$	H_2SO_4	$=$	$NaHSO_4$	$+$	HNO_3
Sodium nitrate		Sulphuric acid		Sodium sulphate		Nitric acid

Experiments with Nitric Acid. After water (H_2O), nitric acid (HNO_3) will dissolve the most substances, and many of them that water (H_2O) has little or no effect upon, such as the heavy metals. Not only this, but nearly all the compounds that are formed by the action of the metals and other substances on nitric acid (HNO_3) will dissolve in water (H_2O).

An Experiment in Spontaneous Combustion. Put a little fuming nitric acid (HNO_3) in a test tube, then wad up some woolen yarn and push it half-way down in the tube. In a little while the yarn will catch fire, and after it has burned up a white substance will be left in the tube. When you make this experiment hold the test tube with your clip in a beaker, so that in case it should break the acid will not do any damage.

The Action of Nitric Acid on Metals. When nitric acid (HNO_3) acts on metals it dissolves them and forms salts that are called nitrates; thus when nitric acid (HNO_3) acts on copper (Cu), copper nitrate ($Cu(NO_3)_2$) is formed; when it acts on silver (Ag), silver nitrate ($AgNO_3$) is formed, and so on, and all these salts will dissolve in water (H_2O).

Cut a silver (Ag) ten-cent piece into bits, put them into a large test tube and add just enough concentrated nitric acid (HNO_3) to cover them. Now hold the test tube over the flame of your alcohol lamp and heat it gently. Colored gases then will be formed, the silver (Ag) will dissolve, and in its place white crystals will be found which are formed of silver (Ag), nitrogen (N), and oxygen (O), and this is silver nitrate ($AgNO_3$).

How the Experiment Works. When nitric acid (HNO_3) acts on the silver (Ag), it gives up part of its hydrogen (H) and oxygen (O) and leaves silver nitrate ($AgNO_3$) and water (H_2O) remaining in the tube. If now you will evaporate the solution, the water (H_2O) will pass off as a vapor and crystals of silver nitrate ($AgNO_3$) alone will remain in the tube.

About Hydrochloric Acid. Since hydrochloric acid ($HCl+H_2O$), or *muriatic acid*, as it is commonly called, or *spirit of salt*, which is its old-time name, is very widely used in the arts, it is a good thing to know something about it. In its pure state it is a colorless gas called hydrogen chloride (HCl), and when this is exposed to the air it fumes, especially in moist air; it has a sour taste which is common to all acids, a strong, pungent odor, and it is very corrosive; further, it has a very great affinity for water (H_2O), and 1 volume of the latter will absorb 500 volumes of the gas.

What we call hydrochloric acid ($HCl + H_2O$) is then hydrogen chloride (HCl) that is dissolved in water (H_2O), and the formula for it is usually given as (HCl) since this is the form of it that is generally known, and the H_2O to show that it is a solution is not considered necessary. When the hydrogen chloride (HCl) and the water (H_2O) are both pure, the hydrochloric acid (HCl) formed of them is colorless, but the acid that is sold for commercial purposes has a yellowish color due to the impurities in it. As hydrochloric acid (HCl) is made of sodium chloride ($NaCl$), that is, common salt, and sulphuric acid (H_2SO_4) it is very cheap, and as it is a most useful acid it is made in large quantities.

HOW TO MAKE HYDROCHLORIC ACID.

To Make Hydrogen Chloride. Put 1 ounce of sodium chloride ($NaCl$) into a flask and then mix ½ ounce of water (H_2O) with 1½ ounces of sulphuric acid (H_2SO_4) in a beaker. Now fit a cork that has a funnel and a delivery tube in it into the flask and let the end of the tube dip into a test tube, as shown in Fig. 114.

This done, pour the solution of sulphuric acid (H_2SO_4) slowly into the flask and heat it gently over the flame of your lamp or burner, and the solution will boil at a great rate. This action causes sodium sulphate (Na_2SO_4) to be formed, while the bubbles that rise up through the solution and break on reaching the surface are the hydrogen chloride gas (HCl), and this passes out of the delivery tube.

To Make Hydrochloric Acid. To change the hydrogen chloride (HCl) into hydrochloric acid (HCl), it is only necessary to fill the test tube with water (H_2O) and let the free end of the delivery tube dip into it while the gas is being generated.

How the Experiment Works. If you have used the right proportions of sulphuric acid (H_2SO_4) and water (H_2O) in making the hydrogen chloride (HCl) and have heated them very gently, all that will be left in the flask will be

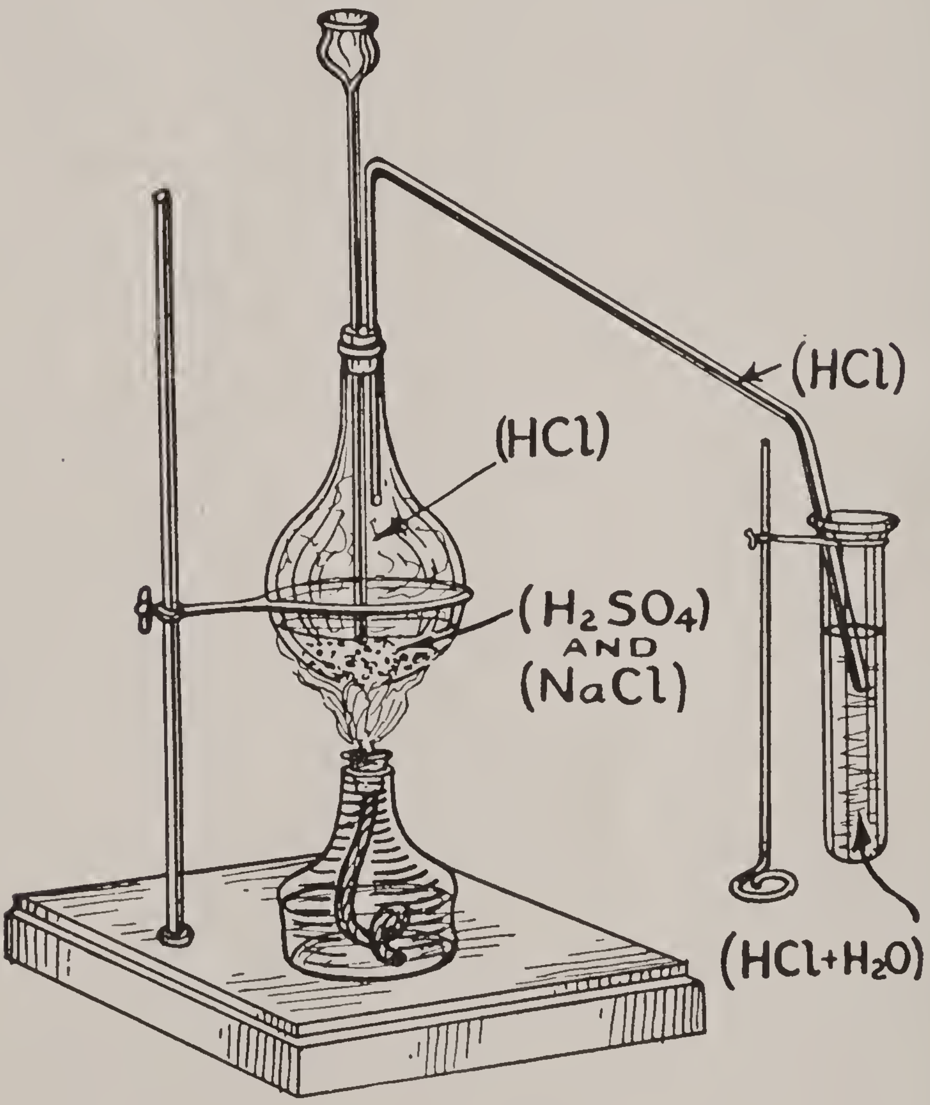

FIG. 114.—How to Make Hydrochloric Acid.

sodium sulphate ($NaHSO_4$), which is a white solid substance. That is to say, the reaction of the sodium chloride ($NaCl$) and the sulphuric acid (H_2SO_4) makes sodium sulphate ($NaHSO_4$) and hydrogen chloride (HCl), and when this is dissolved in water (H_2O) the solution becomes hydrochloric acid (HCl). The following equation shows it in a more simple way.

$NaCl$	+	H_2SO_4	=	$NaSO_4$	+	HCl	↑
Sodium chloride		Sulphuric acid		Sodium sulphate		Hydrochloric acid	A gas

EXPERIMENTS WITH HYDROCHLORIC ACID.

How to Make a Hydrogen Chloride Fountain. Fill a dry flask with hydrogen chloride (HCl) and fit into the neck of it a cork having a glass tube in it, one end of which is drawn out into a nozzle, and also a pipette filled with water (H_2O), the jet of which is closed with a bit of soft wax. Now dip the outside end of the long tube into a beaker of water (H_2O) and you are ready to make the experiment.

Blow the wax out of the end of the pipette by pinching the bulb, and it will send a little stream of water (H_2O) into the flask; the water will absorb so much of the gas that a partial vacuum is formed in the flask. The pressure of the outside air on the water (H_2O) in the beaker will then force it up the tube and into the flask, where it will break into a spray, as shown in Fig. 115.

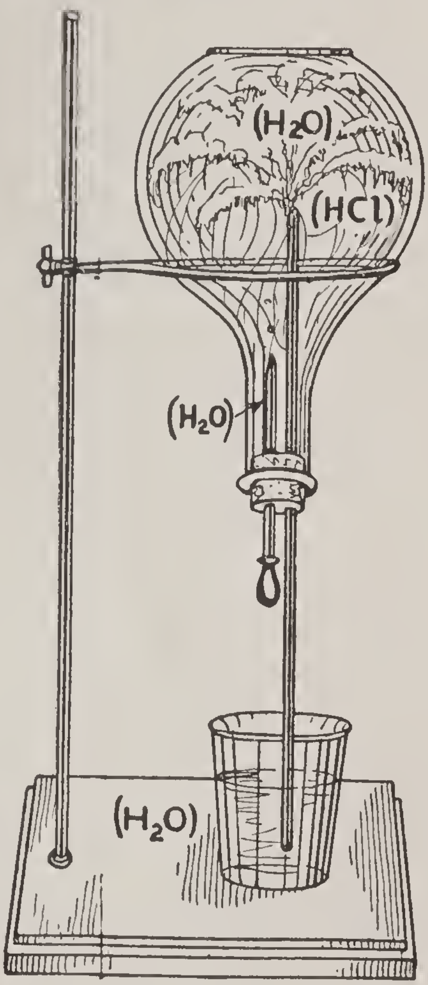

Fig. 115.—A Hydrogen Chloride Fountain.

The Great Smoke Experiment. Fill a wide-mouth bottle with hydrogen chloride (HCl) by inserting the free end of the delivery tube of the generator in it, and then grease a sheet of glass and lay it on top of the bottle. Now fill another bottle with dry ammonia gas (NH_3), as explained in Chapter VI, and close it with a sheet of greased glass also and then set one bottle on the other with their mouths together, as shown in Fig. 116.

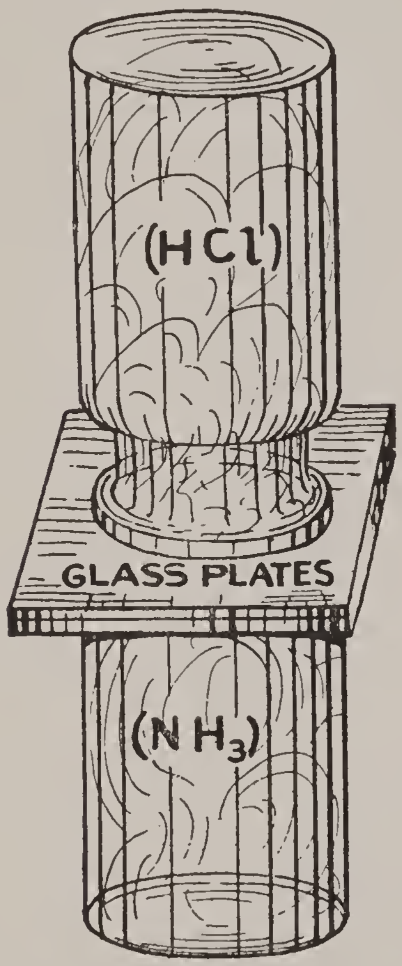

FIG. 116.—The Great Smoke Experiment.

To the onlookers, both bottles will appear to be perfectly empty, but when you pull out the sheets of glass the gases will rush together and form a dense vapor that looks exactly like smoke, and this substance is ammonium chloride (NH_4Cl), or *sal ammoniac*, as it is ordinarily called. A good magical experiment can be performed with a little hydrochloric acid (HCl) and concentrated liquid ammonia (NH_3), and the effect and cause you will find explained in Chapter XIII.

How the Experiment Works. In this experiment the ammonia gas (NH_3) and the hydrogen chloride (HCl) simply combine and make ammonium chloride (NH_4Cl), which is in the form of a fine white powder.

How the Make a Good Soldering Fluid. To make some soldering fluid, cut up a piece of sheet zinc (Zn) into bits, put them into an earthenware bowl, and then pour on enough hydrochloric acid (HCl) to cover them. When the zinc (Zn) is dissolved, the solution makes a good soldering fluid.

How the Experiment Works. When zinc (Zn) is dissolved in more hydrochloric acid (HCl) than is needed, zinc chloride ($ZnCl_2$) is formed, and this remains in the solution. To obtain the zinc chloride ($ZnCl_2$), the solution must be evaporated, and the salt will remain behind. Now by dissolving this in water (H_2O) you can also make a good soldering fluid.

How to Make Imitation Emeralds. Put a few iron (Fe) nails into an earthenware bowl and pour on enough hydrochloric acid (HCl) to cover them. When the nails are dissolved and the solution is yet hot, filter it into a narrow-neck bottle, and solid green crystals that, with a little imagination, look like emeralds, will separate from it, and these are formed of ferrous chloride ($FeCl$).

How the Experiment Works. The reaction in this experiment is as follows: The iron (Fe) and the hydrochloric acid (HCl) form ferrous chloride ($FeCl$) and hydrogen (H), which is set free. When written as an equation you can see the reaction at a glance:

Fe	$+$	HCl	$=$	$FeCl$	$+$	H	$\uparrow$
Iron		Hydrochloric acid		Ferrous chloride		Hydrogen	A gas

How to Make Aqua Regia. The term *aqua regia* comes from two Latin words which mean "water" and "royal", or "royal water". It was so called by the early chemists because it is the only known compound that will dissolve gold (Au) and platinum (Pt), which were and are the *noble metals.*

Aqua regia is therefore a solvent that possesses remarkable properties, and to make it, all you need to do is to mix one part of concentrated nitric acid (HNO_3) and three parts of hydrochloric acid (HCl) in an earthenware bowl.

About Fluorine and Hydrofluoric Acid. Fluorine (F) is a greenish-yellow gas and at room temperature it is the most active element known, for there are very few substances that it will not attack and combine with. The exceptions are oxygen (O), nitrogen (N), chlorine (Cl), platinum (Pt), and those elements of the helium (He) family.

Moreover, it combines usually with all the other elements of its own accord when it is brought into contact with them, and without the application of heat to help the reaction along. The most interesting experiments with fluorine are due to its property of attacking glass (Na_2O, CaO, SiO_2) and other silicate compounds, and the conversion of water (H_2O) into ozone (O_3).

Fluorine (F) is found chiefly in calcium fluoride (CaF_2), which is ordinarily called *fluor-spar*, and also in combination with sodium (Na) and aluminum (Al) in a mineral called cryolite (Na_3AlF_6). Hydrofluoric acid (H_2F_2), is formed when sulphuric acid (H_2SO_4) is made to react with fluor-spar (CaF_2).

How to Etch Glass. To etch on glass (Na_2O, CaO, SiO_2), you have to make some hydrofluoric acid (H_2F_2), and an easy way to do this is to take a sheet of lead (Pb) 5 inches wide and 8 inches long and bend up its edges 1 inch all around to form a tray, as shown in Fig. 117. Now put 2 ounces of powdered calcium fluoride (CaF_2), that is, fluor-spar, into the lead tray and add enough sulphuric acid (H_2SO_4) to make thin paste of it.

This done, melt some paraffin wax, pour it on the sheet of glass which you want to engrave, run it all over the surface and drain it off at one corner, and a thin film of the wax will remain on it. This done, take a darning needle and scratch a name or a picture on the film of wax with the head of it so that the lines go clear through and expose the glass (Na_2O, CaO, SiO_2). Now set the glass (Na_2O, CaO, SiO_2) with the waxed surface down on the tray, light your alcohol lamp and gently heat the paste. The fumes that are given off are hydrogen fluoride (H_2F_2), and this attacks the glass (Na_2O, CaO, SiO_2) and eats it away. After you have exposed it to the fumes for half an hour or

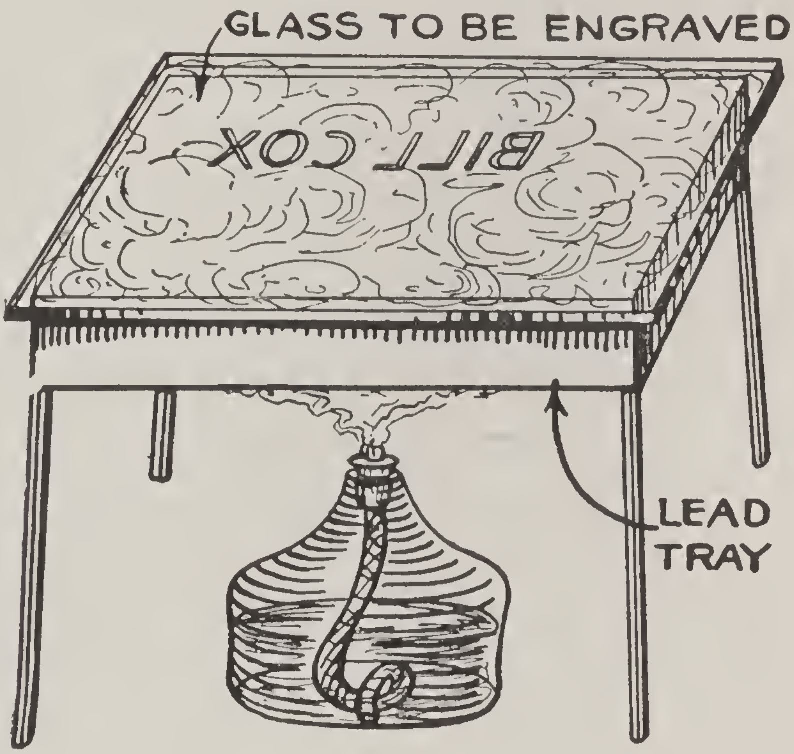

FIG. 117.—How to Etch Glass with Hydrofluoric Acid.

so, wash off the wax with turpentine and you will see the name or picture etched into the glass (Na_2O, CaO, SiO_2).

An Easier Way to Etch on Glass. An easier way to etch on glass (Na_2O, CaO, SiO_2) is to buy some hydrofluoric acid ($H_2F_2 + H_2O$), which is simply hydrogen fluoride (H_2F_2) gas dissolved in water (H_2O). This is sold in rubber bottles, as the acid does not attack rubber (C_5H_8) to any great extent. Coat the glass (Na_2O, CaO, SiO_2) with paraffin wax and scratch a name or draw a picture on and through the film with a needle as before, and then build up a little wax ridge all around the plate, lay it on a table, and with your pipette cover the scratched-in lines with the acid. Let it remain on the plate over night, then wash the wax off with turpentine, and the surface of the exposed parts of the glass (Na_2O, CaO, SiO_2) will be found to be etched away.

How to Change Water into Ozone. If you will turn back to the last part of Chapter II, you will see that ozone (O_3) is produced when electric sparks are made to take place in air. To change the oxygen (O) of water (H_2O) into ozone (O_3), fill a tube with fluorine (F), and as this gas is heavier than the air in the tube, the latter can be held right side up. Now let a single drop of water (H_2O) fall into the tube of fluorine (F) with your pipette, and the oxygen (O) of the water (H_2O) will be turned to ozone (O_3).

CHAPTER VIII.

WHAT BASES AND SALTS ARE.

In the previous chapter I told you about *acids*, and in this chapter I shall tell you about *bases* and *salts*. Now though these three compounds are entirely unlike each other, still they are very closely related, for *salts are formed by the action of acids on bases.* While it is easy to tell an acid if it is a fairly strong one by testing it with litmus paper, some acids are so weak that they have no effect whatever on it, but the fact that they form salts when they are brought into contact with bases shows that they are really acids.

While you may never have heard of bases before, you have often seen them, for they go under the familiar names of lime ($Ca(OH)_2$), soda lye ($NaOH$), and potash lye (KOH), and these compounds all go under the general heading of *alkalis*. These bases are formed by dissolving metals of various kinds in water (H_2O), and the compounds which result easily cut grease and have a very corrosive, or *caustic*, action on the skin and flesh. It is just possible that you did not know there were any metals which would dissolve in water (H_2O); well, there are a few, though these are not at all common in their pure state and, hence, they are seldom seen outside of the laboratory, but they are quite plentiful in the various compounds that nature provides, as you will presently see. These metals are calcium (Ca), sodium (Na), potassium (K), and some others, all of which I shall tell you about in the next chapter.

Finally, a salt is a compound that is formed when an acid and a base combine and during the reaction of which all, or a part, of the hydrogen (H) of the acid is replaced by the atoms of the metal; in this case the hydrogen (H) is set free just as it is when zinc (Zn) or iron (Fe) is dissolved in an acid.

What the Bases Are. As I mentioned above, the three bases that are the most used and, hence, the best known, are caustic lime ($Ca(OH)_2$), caustic soda ($NaOH$), and caustic potash (KOH). Now, as you know, water (H_2O) consists of hydrogen (H) and oxygen (O); calcium (Ca) is a very light metal, and when it is thrown into water (H_2O) it dissolves; in doing so it sets some of the hydrogen (H) of the water (H_2O) free and takes its place, and the compound thus formed is called calcium hydroxide ($Ca(OH)_2$), or caustic lime, or slaked lime, which are its common names.

The word hydroxide means, simply, that the base contains both hydrogen (H) and oxygen (O). Thus calcium hydroxide has the formula ($Ca(OH)_2$), sodium hydroxide has the formula ($NaOH$), and potassium hydroxide the formula (KOH). The most common properties of bases are their alkaline taste and their power to *neutralize* acids, that is, to take away their acid qualities.

What the Salts Are. Now when you bring an acid and a base together, the first result of the action they set up is the formation of water (H_2O), and the second is the formation of a salt. Thus when you make hydrochloric acid

(HCl) act on sodium hydroxide ($NaOH$), water (H_2O) is formed and also a salt that is called sodium chloride ($NaCl$), which is common table salt. Sodium nitrate ($NaNO_3$) is formed by the action of nitric acid (HNO_3) on sodium hydroxide ($NaOH$), potassium chloride (KCl) is formed by the action of hydrochloric acid (HCl) on potassium hydroxide (KOH), potassium nitrate (KNO_3) by the action of nitric acid (HNO_3) on potassium hydroxide (KOH), and so on.

How to Make Calcium Hydroxide. Calcium hydroxide ($Ca(OH)_2$) in spite of its big name is simply *slaked lime.* To make it, you start with calcium oxide (CaO), which is *quicklime,* and this is made by burning limestone ($CaCO_3$) in a kiln. Calcium oxide (CaO) or quicklime, is a very white, porous solid, and when water (H_2O) is poured on it, they unite and a great deal of heat is evolved, and this converts some of the water (H_2O) into steam.

To make a little calcium hydroxide ($Ca(OH)_2$), get some pieces of fresh quicklime, that is calcium oxide (CaO), put them in an earthen bowl and pour a little hot water (H_2O) over them. They will then unite, and the powder which is left is calcium hydroxide ($Ca(OH)_2$), or slaked lime.

How the Experiment Works. The action can be shown thus:

$$\underset{\text{Calcium oxide}}{CaO} + \underset{\text{Water}}{H_2O} = \underset{\text{Calcium hydroxide}}{Ca(OH)_2}$$

How to Make Sodium Hydroxide. Sodium hydroxide ($NaOH$) is *caustic soda,* or *soda lye,* and it is largely used for making soaps from fats. To make sodium hydroxide ($NaOH$), half fill a test tube with water (H_2O), and then dissolve in it ½ teaspoonful of sodium carbonate (Na_2CO_3), the common name of which is *sal soda* and which is often called just *soda* for short, and the same amount of calcium oxide (CaO).

Now hold the test tube over the flame of your lamp or burner for two or three minutes and then put it in the test-tube rack and let it stay there until the solution is perfectly clear. To show that it is an alkali, dip the end of a strip of red litmus paper into it and it will turn blue.

How the Experiment Works. The action that takes place when it is formed is this: the calcium oxide (CaO) and the sodium carbonate (Na_2CO_3) together with the water (H_2O) form calcium carbonate ($CaCO_3$), which settles to the bottom of the tube and sodium hydroxide ($NaOH$) is the clear liquid above it, or:

$$\underset{\text{Calcium oxide}}{CaO} + \underset{\substack{\text{Sodium}\\\text{carbonate}}}{Na_2CO_3} + \underset{\text{Water}}{H_2O} = \underset{\substack{\text{Calcium}\\\text{carbonate}}}{CaCO_3} + \underset{\substack{\text{Sodium}\\\text{hydroxide}}}{NaOH}$$

How to Make Potassium Hydroxide. This hydroxide (KOH), the common names of which are *caustic potash*, *potash lye*, and just *lye* for short, can be used for making hard soap, but as it is more costly than sodium hydroxide ($NaOH$) it is not used by commercial soap-makers. It is, however, a familiar substance in the household and is used for cleaning purposes and making soft soap.

To make potassium hydroxide (KOH), you start with potassium carbonate (KCO) or potash, as it is called, or you can use wood ashes, which contain considerable amounts of it, and, indeed, this used to be the only source from which potash, or lye, was obtained. Dissolve 1 tablespoonful of potassium carbonate (K_2CO_3) or 2 tablespoonfuls of wood ashes in a beaker half full of water (H_2O), and heat this over the flame of your lamp until the solution begins to boil.

Now put in 1 teaspoonful of calcium hydroxide ($Ca(OH)_2$), which is slaked lime that has been made with good quicklime, and keep stirring the mass with an iron rod. This done, let the solution, which is the potassium hydroxide (KOH), get cool, and pour it off very carefully into a bottle, after which it is ready to use for your experiments. Keep it well corked up, or the air will change it back into potassium carbonate (K_2CO_3).

How the Experiment Works. The reaction in this case is that the potassium carbonate (K_2CO_3) and calcium hydroxide ($Ca(OH)_2$) form calcium carbonate ($CaCO_3$) and potassium hydroxide (KOH). As the calcium carbonate ($CaCO_3$) is insoluble, that is, it will not dissolve in the solution, it falls to the bottom, while the potassium carbonate (K_2CO_3) and the calcium hydroxide ($Ca(OH)_2$) are soluble, that is, they dissolve in the solution, and it becomes potassium hydroxide (KOH).

The reaction is more easily understood from the following equation:

$$\underset{\substack{\text{Potassium}\\ \text{carbonate}}}{K_2CO_3} + \underset{\substack{\text{Calcium}\\ \text{hydroxide}}}{Ca(OH)_2} = \underset{\substack{\text{Calcium}\\ \text{carbonate}}}{CaCO_3 \downarrow} + \underset{\substack{\text{Potassium}\\ \text{hydroxide}}}{KOH}$$

Wherever you come to an arrow pointing down in an equation it shows that the compound is precipitated.

EXPERIMENTS WITH HYDROXIDES.

How to Make Mortar. Mix equal parts of calcium hydroxide ($Ca(OH)_2$), that is, *slaked lime*, and *sand*, which latter is formed largely of silicon dioxide (SiO_2), and you will have a sample of mortar. Now let this stand in the air, and the water (H_2O) will soon dry out and calcium carbonate ($CaCO_3$) is slowly formed, and the mass becomes very hard, or sets, as it is called.

How the Experiment Works. First of all, the sand does not act chemically on the lime in any way, but simply serves to give it body and to make it porous. As the water (H_2O) is drying out from the lime, the latter absorbs carbon dioxide (CO_2) from the air; this changes the lime into calcium carbonate ($CaCO_3$), and this with the sand makes a hard, solid mass. The following equation shows the action that takes place:

$Ca(OH)_2$	+	CO_2	=	$CaCO_3$	+	H_2O
Calcium hydroxide		Carbon dioxide		Calcium carbonate		Water (which dries out)

Other Things Made with Lime. Plaster is simply calcium hydroxide ($Ca(OH)_2$), that is, slaked lime mixed with *hair*, which holds it together when the former has changed into calcium carbonate ($CaCO_3$). *Portland cement* is made by heating together slaked lime ($Ca(OH)_2$) and *aluminum silicate* ($Al_2Si_2O_7, H_2O$), which is a hydroxide, and after the mass has been fused by the application of heat it is rolled to a powder. Concrete is simply Portland cement, broken rock, and sand mixed together with water (H_2O). When this sets, a material results that is as hard as stone and that lasts as long.

How to Make Hard Soap. This is an experiment with sodium hydroxide ($NaOH$) and *fat.* The fats used for making soaps are formed of palmitic acid ($C_{16}H_{32}O_2$) or stearic acid ($C_{18}H_{36}O_2$), and *soaps* are the *alkali salts* of the acids that are made when the fats are treated with sodium hydroxide ($NaOH$).

To make a little soap, put a bit of *lard*, which is a fat, about the size of a hazelnut, into a small saucer or, better, a porcelain evaporating-dish, and then pour a teaspoonful of wood alcohol (CH_3OH) over it. This done, dissolve a teaspoonful of sodium hydroxide ($NaOH$) in a like amount of water (H_2O) in a test tube, and then put 10 drops of this solution into the mixture in the dish. Now heat the dish over a low flame until the solution boils and all the alcohol (CH_3OH) has evaporated, which you will know when there is no longer any odor from it; then evaporate the solution slowly until a dry mass results, which will, or should be, soap. If it is not soap, then put a little more alcohol (CH_3OH) and sodium hydroxide ($NaOH$) in the dish and boil it again.

How the Experiment Works. When the sodium hydroxide ($NaOH$) and the stearic acid ($C_{16}H_{32}O_2$), which forms the fat, are heated together, the latter is decomposed and forms glycerine ($C_3H_5(OH)_3$) and sodium stearate ($NaC_{18}H_{35}O_2$), and these together are the *soap.* The alcohol (CH_3OH) simply helps the reaction along and does not enter chemically into the process. When the reaction takes place the fat is said to be *saponified* and the process is called *saponification.* The following equation shows the reaction more clearly:

$$\underset{\substack{\text{Stearic acid}\\\text{in fats}}}{C_{18}H_{36}O_2} + \underset{\substack{\text{Sodium}\\\text{hydroxide}}}{NaOH} = \underset{\text{Glycerine}}{C_3H_5(OH)_3} + \underset{\substack{\text{Sodium stearate}\\\text{Soap}}}{NaC_{18}H_{35}O_2}$$

How to Make Soft Soap. To make *soft soap*, you need only to mix a little lard with some potassium hydroxide (KOH) and boil them, and they will saponify. Potassium hydroxide (KOH) is not used in making soap commercially, because sodium hydroxide ($NaOH$) is cheaper and does the work just as well or better.

How Soap-and-Water Cleans. When you mix soap and soft water (H_2O) they make a soap solution and this cleans the stains from the goods washed in them by separating the *vegetable* and *animal* oils and grease and washing out the dust and dirt that is in them. The first thing that happens when goods are washed in a soap solution is that the soap breaks up the oils and grease into little separate particles, and then forms a film around each one, like the sugar coating on a pill. When this is done, the water (H_2O) easily washes them away. If the goods have *mineral oils* in them, then they must be *dry cleansed*, that is, cleaned with *benzine* or *gasoline*, which dissolves the oils.

When the clothes are *dirty*, it means simply that they have a large amount of particles of carbon (C) on and in them, and much of this is soot. When the soap solution comes into close contact with them, which it does when the clothes are boiled and rubbed, the particles of dirt are easily washed away by rinsing the clothes in water (H_2O).

HOW TO MAKE VARIOUS SALTS.

How to Make Sodium Chloride. Since you can buy a pound of sodium chloride ($NaCl$), that is, common table salt, for a few cents it is not good economy to make it, but it is well worth your while to do so for the experience it will give you. To make a little, dissolve a tablespoonful of sodium hydroxide ($NaOH$), that is, caustic soda, in a small beaker half full of water (H_2O).

Now fill a test tube one-fourth full of water (H_2O) and add an equal amount of hydrochloric acid (HCl). This done, pour the dilute acid, a very little at a time, into the solution of caustic soda ($NaOH$) and keep testing it with blue litmus paper right along; the instant it changes the paper from blue to red add a few drops of caustic soda ($NaOH$), and it will be neutral, that is, it will be neither acid nor alkali and hence, it will not change blue litmus paper red nor red paper blue.

When this point is reached, pour the solution from the beaker into a porcelain dish and evaporate it over a water bath until there is nothing left but a white residue which is neither an *alkali* nor an *acid* but a *salt*, and this is sodium chloride ($NaCl$).

How the Experiment Works. The reaction that takes place is clearly shown by the following equation:

$$\underset{\text{Hydrochloric acid}}{HCl} + \underset{\substack{\text{Sodium}\\ \text{hydroxide}}}{NaOH} = \underset{\substack{\text{Sodium}\\ \text{chloride}}}{NaCl} + \underset{\text{Water}}{H_2O}$$

How to Make Sodium Sulphate. Put a little sodium chloride ($NaCl$), into a test tube and pour on enough sulphuric acid (H_2SO_4) to cover it. Now hold it over the flame of your lamp or burner, and very soon hydrochloric acid gas (HCl) and sodium sulphate (Na_2SO_4), the common name of which is *Glauber's salt*, will be formed; the gas will, of course, escape and the salt will be left behind.

How the Experiment Works. The following equation shows the reaction:

$NaCl$	$+$	H_2SO_4	$=$	Na_2SO_4	$+$	$HCl \uparrow$
Sodium chloride		Sulphuric acid		Sodium sulphate		Hydrochloric acid gas

How to Make Sodium Nitrate. To make some sodium nitrate ($NaNO_3$), which is commonly called *Chili saltpetre*, proceed exactly as described in the above experiment, but use nitric acid (HNO_3) and caustic soda ($NaOH$) instead. The reaction that takes place gives sodium nitrate ($NaNO_3$) and water (H_2O).

How the Experiment Works. The following equation shows what takes place:

HNO_3	$+$	$NaOH$	$=$	$NaNO_3$	$+$	H_2O
Nitric acid		Caustic soda		Sodium nitrate		Water

The solution must be evaporated, and when the water (H_2O) has passed off, the sodium nitrate ($NaNO_3$) alone will remain.

How to Make Potassium Chloride. This salt ($KClO_3$) is a white substance in the form of crystals, and the larger part of it comes from the great potash beds at Stassfurt, Germany. Giant kelp, or seaweed, as it is commonly called, contains about 9 per cent of it, and from this it is now extracted to some extent commercially.

To make a little potassium chloride ($KClO_3$), follow the directions given for making sodium chloride ($NaCl$), but use hydrochloric acid (HCl) and potassium hydroxide (KOH), that is, caustic potash for the base.

How the Experiment Works. The reaction will give you potassium chloride ($KClO_3$) and water (H_2O), thus:

HCl	+	KOH	=	$KClO_3$	+	H_2O
Hydrochloric acid		Potassium hydroxide		Potassium chloride		Water

Evaporate the solution that remains after heating the acid and the base together and the potassium chloride (KCl) will remain behind in the form of a crystalline salt.

How to Make Potassium Nitrate. Potassium nitrate (KNO_3) or just *saltpetre*, as it is usually called for short, is also a white crystalline salt and is chiefly used for making black gunpowder. Now while sodium nitrate ($NaNO_3$) is cheaper than potassium nitrate (KNO_3), the former cannot be used in making gunpowder, as it is *deliquesent*, that is, it absorbs moisture from the air, while the latter does not. To make potassium nitrate (KNO_3) on a commercial scale, potassium chloride ($KClO_3$) from the Stassfurt potash beds is added to a hot solution of sodium nitrate ($NaNO_3$).

You can make a sample of potassium nitrate (KNO_3) by using the same process as described for the foregoing salts, but in this case add nitric acid (HNO_3) to potassium hydroxide (KOH), and potassium nitrate (KNO_3) and water (H_2O) will result.

How the Experiment Works. The reaction is this:

HNO_3	+	KOH	=	KNO_3	+	H_2O
Nitric acid		Potassium hydroxide		Potassium nitrate		Water

The water (H_2O) is then evaporated as explained before, and the potassium nitrate (KNO_3) salts are left behind.

CHAPTER IX.

THE MYSTIC METALS

THEIR ALLOYS AND AMALGAMS.

All the elements can be classified under two general headings, namely, those that are *metals* and those that are *non-metals*. Now while it is easy to tell a metal from a non-metal when you see it, it is not at all easy to define the difference; it will, however, serve the purpose to say that a metal is an element which is opaque, has a metallic lustre, is a good conductor of electricity, and, finally, and most important of all, is able to take the place of the hydrogen (H) in an acid and to form a salt.

There are two ways in which metals occur in nature, and these are in a *native*, or *free*, state, that is, they are found in a pure form, and when they are mixed, or combined, with other substances, and when these are hard they are called *ores*. Copper (Cu), lead (Pb), silver (Ag), gold (Au), platinum (Pt), and some other metals are found in a free state.

Of the ores, there are several chief kinds, and these are the *oxides*, the *sulphides*, and the *carbonates*.

Among the oxides are those of iron (FeO) and (Fe_2O_3), zinc (ZnO), tin (SnO_2), or *tin stone*, as this ore is called, copper (Cu_2O), which is called *ruby copper*, etc. Among the sulphides are those of iron (FeS_2), or *iron pyrites*, nickel

(NiS), cobalt ($CoAsS$) or *cobaltite*, as this ore is called, antimony (Sb_2S_3), or *stibnite*, lead (PbS), etc. Finally, among the carbonates are those of iron ($FeCO_3$), lead ($PbCO_3$), zinc ($ZnCO_3$) and copper ($Cu_2(OH)_2CO_3$) or *malachite*.

The Activity of the Metals. The power of a metal to displace hydrogen (H) in dilute acids and water (H_2O) is called its *activity*. Now the most active metal is potassium (K), as the table given below shows, while lead (Pb) is the least active, and all those below the Zero or hydrogen(H) line will not dissolve in water (H_2O) or weak acids, and, hence, will not replace the hydrogen (H) in them at all, and they are, therefore, called the *inactive* metals. These metals can, however, be dissolved in strong acids.

TABLE OF ACTIVITIES.

13\. Potassium (K)
12\. Sodium (Na)
11\. Lithium (Li)
10\. Calcium (Ca)
9\. Magnesium (Mg)
8\. Aluminum (Al)
7\. Manganese (Mn)
6\. Zinc (Zn)
5\. Chromium (Cr)
4\. Iron (Fe)
3\. Nickel (Ni)
2\. Tin (Sn)
1\. Lead (Pb)
0\. Hydrogen (H)
1\. Copper (Cu)
2\. Bismuth (Bi)
3\. Antimony (Sb)
4\. Mercury (Hg)
5\. Silver (Ag)
6\. Platinum (Pt)
7\. Gold (Au)

There are many other metals than those given in the preceding table, but these are the best known, and their distinguishing features, various uses, and experiments with them will follow in the above order.

Potassium, the Softest Metal. The Latin word for potassium (K) is *kalium*, and since P is the symbol used for phosphorus (P), which was a much earlier known element, the letter K was chosen for potassium (K). This metal was discovered in 1807 by Sir Humphrey Davy, who made it by passing a current of electricity through some potassium hydroxide (KOH), causing minute drops of the pure metal to be formed on the negative wire. In the early days, potassium hydroxide (KOH) was obtained from wood-ashes, and when these were leached, boiled, and evaporated, the remaining substance was called *pot-ashes*, then just *potash*, and from this we get the name potassium (K).

Potassium (K) is a silvery-white metal with a bright metallic lustre and so soft that you can knead it with your fingers at room temperature, just as you would a piece of wax. Owing to its great activity when it comes into contact with water (H_2O), it must be kept in oil so that the moisture of the air cannot get to it. It is so light that it floats on water (H_2O), and it melts at a much lower temperature than that at which water (H_2O) boils.

Compounds of Potassium. While potassium (K) is of little use by itself, the compounds made with it are of great value. Its chief compounds are potassium iodide (KI), which is used for testing starch, in medicine, and in photography; potassium hydroxide (KOH), which is used largely for making other compounds of potassium (K); potassium nitrate (KNO_3), which is used for preserving meats and for making gunpowder and fireworks; potassium chlorate ($KClO_3$), which is used in making oxygen (O) and in medicine, and potassium carbonate (K_2CO_3), which is used as a fertilizer.

An Experiment with Potassium. Take a piece of potassium (K) out of the bottle of oil with your tweezers, cut off a piece the size of a pea,and drop it into a bowl of water (H_2O); instantly the hydrogen (H) of the latter will be set free and the heat produced will ignite the potassium (K); the gas will then force the burning metal through the water (H_2O), as shown in Fig. 118, and as it darts along it will make explosive sounds like a bunch of miniature firecrackers going off.

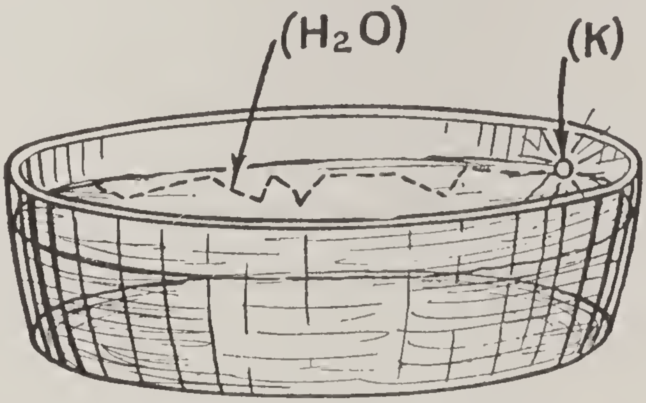

FIG. 118.—The Reaction of Potassium on Water.

How the Experiment Works. When you drop the potassium (K) on the water (H_2O), the chemical action sets the hydrogen (H) free so fast that it develops enough heat to ignite both the gas and the metal, and the mechanical reaction between the escaping gas and the metal forces the latter along on the water (H_2O). The chemical reaction that takes place between the potassium (K) and the water

(H_2O) forms potassium hydroxide (KOH) and hydrogen (H) thus:

$$\underset{\text{Potassium}}{K} + \underset{\text{Water}}{H_2O} = \underset{\text{Potassium hydroxide}}{KOH} + \underset{\text{Hydrogen}}{H \uparrow}$$

Sodium, Another Alkali Metal. Since S is the symbol of sulphur (S) one of the earliest known elements, the letters Na, are used for sodium (Na), because in Latin the latter was called *natrium*, which means *soda*. This metal was also discovered by Davy, who made it in 1808 by bringing a pair of wires from a battery into contact with sodium hydroxide ($NaOH$), that is, caustic soda. Sodium (Na) is a soft, shining, silvery metal, and it behaves very much like potassium (K) when it is dropped into cold water, but since it is not so active as the latter metal it does not produce enough heat to ignite the hydrogen (H) which is set free around it.

Compounds of Sodium. The pure metal is chiefly of interest for experimental work, but it is widely employed by nature and the chemist in making various salts, and in combination with carbon (C) compounds it is used for both dyes and drugs. The chief compounds in which it occurs are sodium chloride ($NaCl$), that is, common table salt; sodium nitrate ($NaNO_3$), which is the starting-point in making potassium nitrate (KNO_3), or saltpetre, and of nitric acid (HNO_3); sodium carbonate (Na_2CO_3), which in turn, is used for making *sodium bicarbonate* ($NaHCO_3$), or baking soda, etc.

An Experiment with Sodium. Drop a small piece of sodium (Na) into a dish of cold water (H_2O) and watch its action. Having made the above experiment, thicken the water (H_2O) with a little starch ($C_6H_{10}O_5$), which will cause the heat developed to be concentrated, and then the hydrogen (H) will ignite and the sodium (Na) will burn with a brilliant yellow color.

How the Experiment Works. The reaction that takes place when sodium (Na) comes in contact with water (H_2O) is that they form sodium hydroxide ($NaOH$) and hydrogen (H). The following equation shows at a glance what takes place:

$$\underset{\text{Sodium}}{Na} + \underset{\text{Water}}{H_2O} = \underset{\text{Sodium hydroxide}}{NaOH} + \underset{\text{Hydrogen}}{H\uparrow}$$

Lithium, the Lightest Metal. This metal gets its name from the Greek word *lithium*, which means *stone*. Now it happens that while lithium (Li) is obtained from stone-like minerals, such as *lapidolite*, which is a kind of mica and is, therefore, quite heavy, the metal itself is the lightest of them all, and what is more, it is the lightest solid known; in fact, it is so light that it not only floats on water but on the oil in which it is kept.

Traces of lithium (Li) are found in the water of various mineral springs, in the soil, in the ashes of tobacco, and in beets. It is made by passing a current of electricity through fused lithium chloride ($LiCl$). When brought into contact with water (H_2O), it acts like potassium (K) and sodium (Na), except that it is not so active, but different from these metals, it is quite hard.

Compounds of Lithium. Lithium (Li) forms compounds with hydrogen (H), nitrogen (N), and oxygen (O), but unlike those of the other alkali metals the hydroxide ($LiOH$) formed of it does not dissolve easily in water (H_2O).

An Experiment with Lithium. Drop a piece of lithium (Li) into a test tube half full of cold water (H_2O), and then drop a piece into a like amount of warm water (H_2O), and you will see that it combines with the latter very much faster than with the former. When the lithium is dissolved, the compound that is formed is lithium hydroxide ($LiOH$).

Calcium, the Fourth Alkaline Metal. The Latin word for *lime* is *calx*, and from this we get the word calcium (Ca), which is the chemical name for lime. The metal in its pure state looks like silver (Ag), and melts when brought to a cherry-red heat. It is not quite as soft as lead (Pb) and, like the latter, it can be cut, drawn, and rolled.

Calcium (Ca) is never found in a free state, but is very plentiful in different compounds, especially in calcium carbonate ($CaCO_3$), of which chalk, marble, and limestone are formed; it is also found in calcium sulphate ($CaSO_4$), which is gypsum, and as calcium phosphate ($Ca_3(PO_4)_2$), in the minerals *apatite* and *phosphorite*, and in fluoride (CaF_2), that is, *fluor-spar*. It is also found in plants, in the bones of animals, and in sea-shells, while coral and pearls are made of it.

Compounds of Calcium. Like the other alkaline metals, calcium (Ca) is of but little use in its pure state, but its compounds are very valuable. Chief among these are calcium chloride ($CaCl_2$), calcium oxide (CaO), that is, quicklime, calcium hydroxide ($Ca(OH)_2$), or slaked lime, calcium carbonate ($CaCO_3$), which, as I have said so many times before, is limestone.

Now limestone ($CaCO_3$) will dissolve in water (H_2O) that has carbon dioxide (CO_2) in it, which then acts like a weak acid and, indeed, it is called carbonic acid (H_2CO_3). When the water (H_2O) of a river runs over limestone ($CaCO_3$), it dissolves it and sometimes forms a cave, or a cavern; on the other hand, when water (H_2O) has calcium bicarbonate ($Ca(HCO_3)_2$) in it and it seeps drop by drop through the ceiling of the cave, or cavern; it loses its carbon dioxide (CO_2) and the calcium bicarbonate ($Ca(HCO_3)_2$) gets hard and forms *stalactites*, the name given to needles which hang from the ceiling, and *stalagmites*, needles which rise from the floor.

Experiments with Calcium. Take a piece of calcium (Ca), stand about 3 feet from a clean brick wall and throw it as hard as you can against it, and the metal will ignite and burn with a brilliant white flame.

Spread some calcium sulphide (CaS), which is a compound made by heating pulverized calcium sulphate ($CaSO_4$) and charcoal (C) together, on a sheet of paper and lay it where the sunlight will strike it. After you have thus exposed it for half an hour or so, take it into a dark room and it will shine with a cold light. Calcium sulphide (CaS) and barium sulphide (BaS), which also shines in the dark, are used for making *luminous paint.*

Magnesium, the Metal that Burns. Magnesium (Mg) gets its name from Magnesia, a district in Asia Minor, where magnesium carbonate ($MgCO_3$), or magnetite, as it is called, was first found. When Davy was experimenting with the action of an electric current on various substances, he discovered that magnesium carbonate ($MgCO_3$) was a compound of a metal with carbon (C) and oxygen (O); then in 1830, Bussy, a French chemist, separated the metal. This he did by treating magnesium chloride ($MgCl_2(H_2O)$), a salt that is found in salt deposits, with potassium chloride (KCl), causing the potassium (K) to combine with the chlorine (Cl) and leave the magnesium (Mg) in a pure state.

It is now obtained by passing a current of electricity through melted magnetitie ($MgCO_3$), potassium chloride (KCl), and sodium chloride ($NaCl$). It is a silvery-white metal, quite soft and very light and brittle. Magnesium (Mg) is used chiefly for making flash-lights for taking photographs, as it is rich in violet waves, which act strongly on a photographic plate, for making fireworks and signal lights, and for making silver polish and tooth powders.

Compounds of Magnesium. Chief among the compounds of magnesium (Mg) are magnesium oxide (MgO) and the hydroxide ($Mg(OH)_2$). The former is made by heating magnesium carbonate ($MgCO_3$), when the product is called *calcined magnesia*, and as even the highest temperatures will not affect this, it is used for making crucibles and lining furnaces. When water (H_2O) is poured over magnesium oxide (MgO), they combine slowly and form magnesium hydroxide ($Mg(OH)_2$). When magnesium (Mg) is combined with oxygen (O), carbon (C), and hydrogen (H), they form a *magnesium alba*, as it is commonly called, and this is largely used in medicines and cosmetics.

Experiments with Magnesium. Magnesium (Mg) can be bought in the form of thin ribbon coiled up, or in a powder. Take a piece of the ribbon and light the end of it with a match and it will give out a light of dazzling brightness, as shown in Fig. 119. The greyish powder that remains as an ash after it is burned is magnesium oxide (MgO). Mix a teaspoonful of powdered magnesium (Mg) with 1½ teaspoonfuls of potassium chlorate ($KClO_3$) and you will have a flash-light powder.

Aluminum, the Lightest Common Metal. Aluminum (Al) gets its name from *alumen*, which means *clay*. After oxygen (O) and silicon (Si), aluminum (Al) is the most plentiful element we have, and all kinds of clays and micas contain it. Davy discovered the metal in 1808, and he first obtained it from aluminum sulphate ($Al_2(SO_4)_3H_2O$), or *alum*, which is its common name, and this is made by treating pure clay with sulphuric acid (H_2SO_4). The metal is now made on a large scale by passing an electric current through aluminum oxide (Al_2O_2), which is dissolved in melted cryolite ($NaF+AlF_3$), a mineral of Greenland, that is formed of sodium fluoride (NaF) and aluminum fluoride (AlF_3). Aluminum (Al) is a whitish metal, the color of which lies between that of tin (Sn) and zinc (Zn), it is light, strong, and tough, weighs about 2½ times as much as water (H_2O), and is both malleable and ductile. It is very resonant, and when a rod of it fixed at one end is struck with a piece of wood it gives a swelling musical sound.

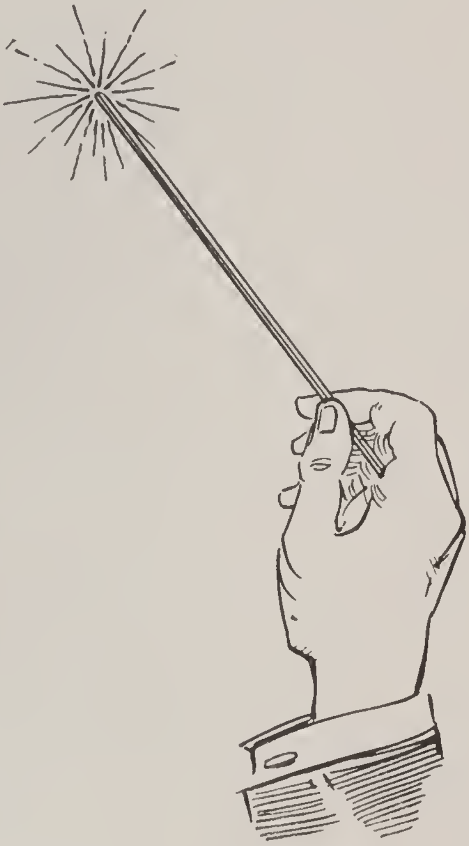

FIG. 119.—Magnesium Burning in Air.

Aluminum (Al) is largely used instead of copper (Cu) for electric transmission lines, for making utensils, in making pure steel, and for thermit, a high-temperature compound. The steel-makers put 1 part of aluminum (Al) in with 1000 parts of steel, and when the metals are melted the gases combine with the aluminum (Al), so that when the steel is cast it is free from blow-holes. The chief compounds of aluminum are aluminum oxide (Al_2O_3); aluminum hydroxide (AlO_3H_3); aluminum chloride ($AlCl_3$); aluminum silicate ($AlKSi_3O_8$), which when mixed with water (H_2O) is clay; and the alums, which are compounds formed of potassium (K) and aluminum (Al), called potassium alum ($AlK(SO_4)_2+12\,H_2O$); and sodium (Na) and aluminum (Al), called sodium alum ($AlNa(SO_4)_2+12\,H_2O$). When aluminum (Al) burns, it produces a very high temperature — about 3000 degrees Fahrenheit — and this will melt most metals. This fact is taken advantage of in thermit, which is the trade-name for a process. The following experiment shows how it works.

An Experiment with Aluminum. Mix a tablespoonful of powdered aluminum (Al) and a like amount of iron oxide (Fe_2O_3) in a small sand crucible; now push a piece of magnesium (Mg) wire down into the mixture and light it, as shown in Fig. 120, and the aluminum will take fire. The burning aluminum (Al) melts the iron (Fe) in the iron oxide (Fe_2O_3), and when the mass is cold you will find a button of pure iron (Fe) in the bottom of the crucible. This experiment must be done out-of-doors.

How the Experiment Works. The reason magnesium (Mg) wire is used to ignite the aluminum (Al) powder is because it takes a hotter flame to fire the latter than a burning splint of wood can produce. The aluminum (Al) gets oxygen (O) in which to burn from the iron oxide (Fe_2O_3), and these combine to form aluminum oxide (Al_2O_3), while pure iron (Fe) remains in the bottom of the crucible. The reaction is this:

Al	+	Fe_2O_3	=	Al_2O_3	+	Fe
Aluminum		Iron oxide		Aluminum oxide		Iron

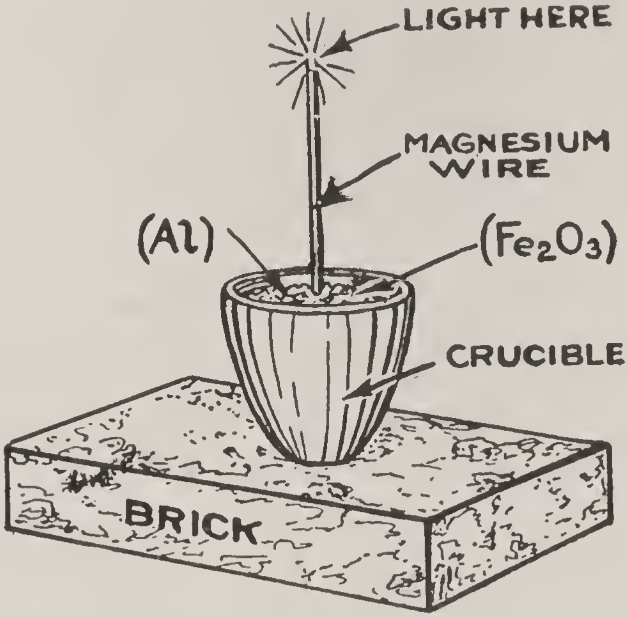

FIG. 120.—Making Iron by the Thermit Process.

Manganese, the Hardening Metal. The Latin word for magnet is *magnes* and manganese (Mn), was named after it, not only because the ore in which it was found looked something like the natural magnet, or *lodestone*, as it is called, but also because it is slightly magnetic. This metal was discovered by Gahn in 1775 in an ore that Scheele named *manganese.*

The metal does not occur free in nature but is found in large quantities in a mineral called *pyrolusite*, which is crude black manganese dioxide (MnO_2). It can be extracted by mixing this mineral and aluminum (Al), both in powdered form, in a crucible and igniting them, as explained under the heading of an "Experiment with Aluminum." It is a heavy, hard, and brittle metal of a grey color.

Compounds of Manganese. When manganese (Mn) combines with oxygen (O) it forms several compounds, the most common of which are manganese oxide (MnO) and manganese dioxide (MnO_2). It also forms several compounds when it combines with potassium (K), and among these are potassium manganate (K_2MnO_4) and potassium permanganate ($KMnO_4$). When mixed with steel, it forms an alloy of exceeding hardness, and this will be described farther along.

An Experiment with Manganese. Dissolve ½ teaspoonful of ferrous sulphate ($FeSO_4$), that is, *copperas*, or *green vitriol*, in a test tube half full of water (H_2O) and add a couple of drops of sulphuric acid (H_2SO_4) to it. This done, dissolve ½ teaspoonful of potassium permanganate ($KMnO_4$) in a test tube half full of water (H_2O) and you will have a solution of a deep purplish-red color. Now with your pipette add a drop at a time of the former to the latter, and the color will disappear.

Zinc, the Electric Metal. While zinc (Zn) is the negative element that is used in all primary battery cells for generating a current of electricity, it finds a wider application in the building industries, since it does not rust like iron (Fe), and in making brass and other alloys. It is found in various ores, and among these are *sphalerite*, or *zinc-blende*, (from the German *blenden*, meaning *to dazzle*), which is zinc sulphide (ZnS), and also in *smithsonite* ($ZnCO_3$).

To separate the zinc from these compounds, in the first case, the sphalerite is crushed and then roasted, causing the sulphur (S) to pass off, and in the second case the smithsonite is powdered and mixed with coal and then heated, which drives off the (CO_3). The pure metal is of a bluish-white color and can be easily rolled into sheets when it is heated to 150 degrees Fahrenheit, but it is brittle when heated either above or below this point. Zinc oxide (ZnO) or *zinc white*, as it is commonly called by painters, zinc sulphate ($ZnSO_4+H_2O$), or *white vitriol*, to give it its common name, and zinc chloride ($ZnCl_2$) are the chief compounds formed of zinc (Zn).

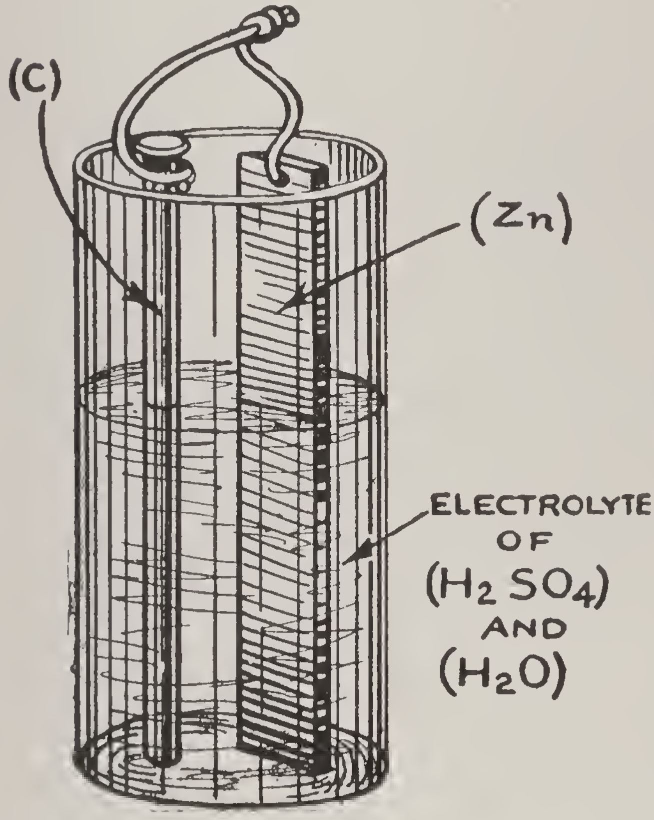

FIG. 121.—A Simple Electric Cell.

An Experiment with Zinc. Make a solution, called an *electrolyte*, by adding 1 fluid ounce of sulphuric acid (H_2SO_4) to a beaker three-fourths full of water (H_2O); put in a strip of zinc (Zn) and a rod of carbon (C), to one end of each of which you have fastened a copper wire, and you will have a simple battery cell. If now you will wind one of the wires around a nail and connect the end of it with the end of the other battery wire, as shown in Fig. 121, the nail will become a magnet, and when this is the case you will know that a current is flowing in the wire.

Chromium, the Color-Making Metal. We get the word *color* from the Greek root *chroma*, and because this metal gives variously colored compounds it is called chromium (Cr). While it is not a very well known metal, still it was discovered over a hundred years ago. It is found in a mineral called *chromite*, which is ferrous chromite ($Fe(CrO_2)_2$), and in *chrocoisite*, which is lead chromate ($PbCrO_4$). The metal is easily extracted from this mineral by the *thermit process*. (See Aluminum). Like manganese (Mn), chromium (Cr) is used for hardening steel (Fe), and alloys containing it will be described presently. The chief compounds of chromium (Cr) are potassium chromate (K_2CrO_4) and potassium dichromate ($K_2Cr_2O_7$).

Experiments with Chromium. Potassium dichromate ($K_2Cr_2O_7$) is the most common compound of chromium (Cr). It is made by dissolving chromium (Cr) in nitric acid (HNO_3). Heat a beaker half full of water (H_2O) until it boils, then remove it from the flame and stir in a teaspoonful of potassium dichromate ($K_2Cr_2O_7$); when it is cold, add 1 ounce of sulphuric acid (H_2SO_4) and stir the solution well with a glass rod. This done, pour it into your porcelain evaporating-dish and cover it with a board, let it stand for several hours, and dark red crystals will be formed in it. These are chromic acid (H_2CrO_4).

Pour off the red solution, then lay the crystals on a piece of flower pot, or other unglazed pottery, and cover them with a glass jar so that the air can not get to them, and in a couple of days they will be quite dry. Put a teaspoonful of dry chromic acid (H_2CrO_4) in your porcelain evaporating-dish and add a few drops of 95 per cent ethyl alcohol (C_2H_5OH), which is grain alcohol, and the latter will immediately burst into flames, as shown in Fig. 122.

Iron, the Most Useful Metal. Iron (*Fe*) is such a common metal that it needs but little description. It is seldom found in pure state but it is very plentiful in the ores of *magnetite* (FeO_4), *hematite* (Fe_2O_3), *siderite* ($FeCO_3$), and pyrites (FeS_2). To get the iron (*Fe*) out of these ores, they are mixed with coke (*C*) in a blast furnace and the coke (*C*) is then burned in a blast of air. The oxide or sulphur (*S*) in them then combines with the gases of the air, and the iron (*Fe*) melts and flows to the bottom of the furnace.

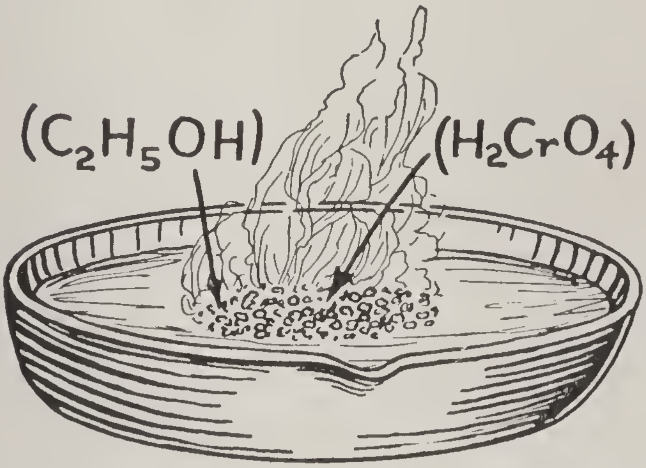

FIG. 122—Chromium Crystals and Alcohol Bursting into Flames.

When other metals and substances are mixed with iron (*Fe*) it takes on new properties, and these alloys will be described farther on. There are numerous iron (*Fe*) chemical compounds, and among them are ferrous carbonate ($FeCO_3$), ferrous sulphate ($FeSO_4$), and ferrous sulphide (*FeS*); also ferric chloride ($FeCl_2$) and ferric oxide (FeO_3), which latter is *iron rust.*

An Experiment with Iron. To make a little ferrous sulphide (*FeS*), dissolve ½ teaspoonful of ferric ammonium sulphate ($(NH_4)_2SO_4 + Fe_2(SO_4)_3 + 24H_2O$) in a test tube half full of water (H_2O). Now put ½ teaspoonful of sulphur (S) and a bit of paraffin wax the size of a pea into another test tube; fit a cork that has a bent delivery tube in it into the mouth of the latter, and place the end of this

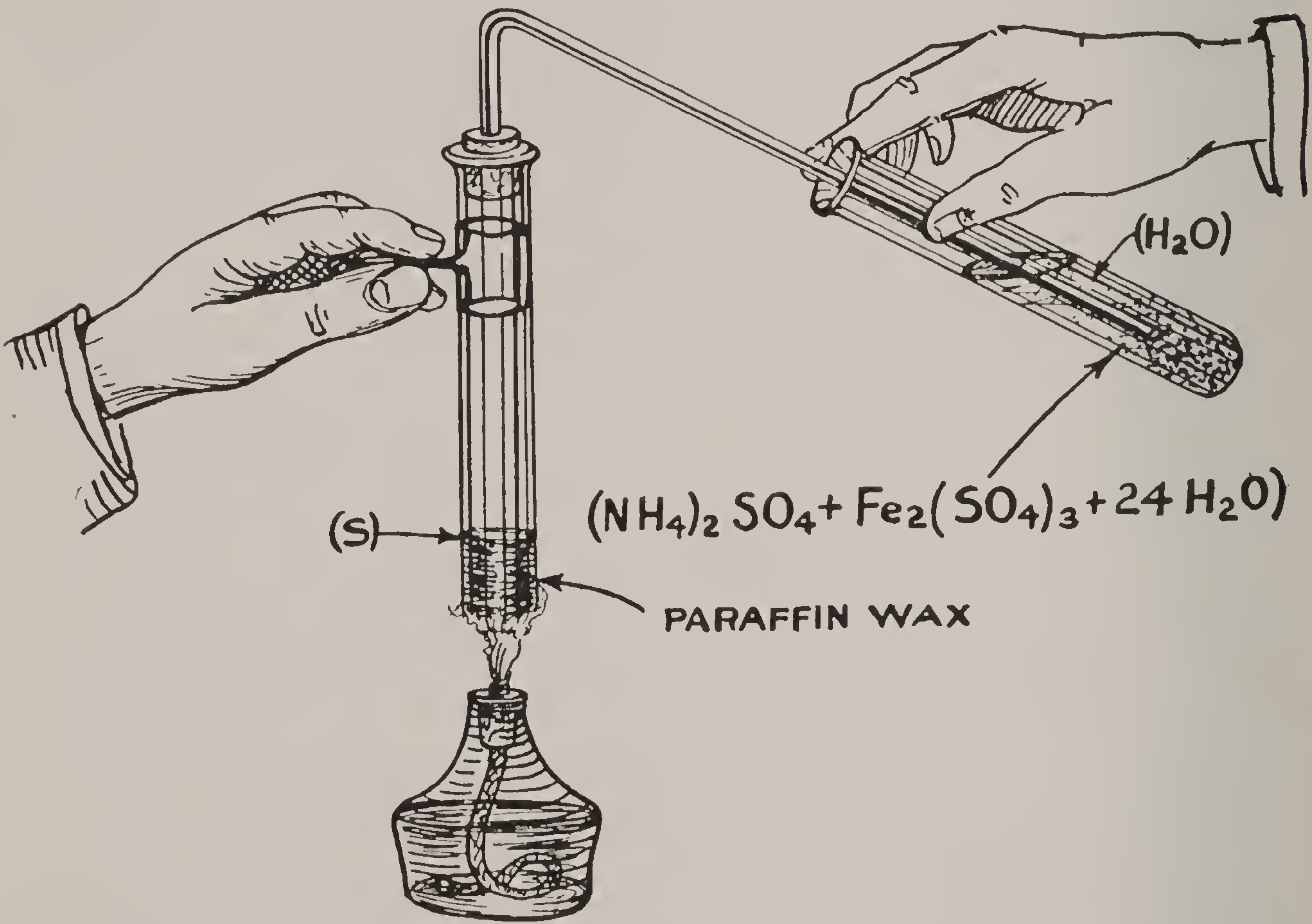

Fig. 123.—Making Ferric Sulphide.

in the test tube containing the ferric ammonium sulphate (see formula above) solution, as shown in Fig. 123. Finally, melt the paraffin and sulphur (S) over the flame of your lamp or burner, and let the hydrogen sulphide (H_2S), which is a gas and smells like rotten eggs, pass through it for 3 or 4 minutes. Soon there will be a black precipitate formed, and this is ferrous sulphide (FeS).

Nickel, the Non-Rusting Metal. Nickel (Ni) is a white metal with a slightly yellowish tinge; it is very hard, has a high melting point, and takes a fine polish. It is never found free except in meteorites, but it is found in combination with arsenic (As), and is also obtained from a mineral called pentlandite ($(Ni,Cu,Fe)S$). To extract the nickel (Ni) from its ore, the latter is roasted, which drives off the sulphur (S), and then smelted to separate it from the copper (Cu) and the iron (Fe). As it rusts very slowly in moist air and because it takes a beautiful silvery polish, it is largely used in electroplating iron (Fe) and steel (Fe,C) articles of all kinds.

There are many compounds of this metal, including the chloride ($NiCl_2,H_2O$), the sulphate ($NiSO_4,H_2O$), and the oxide (Ni_2O_3). Nickel ammonium sulphate ($(NH_4)_2SO_4,NiSO_4,H_2O$) is used for making nickel sulphide (NiS), nickel carbonate ($NiCO_3$), nickel tetraborate ($NiB_4(OH)_3$), and other nickel compounds.

How to Nickel-Plate a Coin. To electroplate a copper coin with nickel (Ni), dissolve a teaspoonful of nickel ammonium sulphate (see formula above) in a test tube two-thirds full of boiling water (H_2O). Now wrap a thin copper wire around the coin and dip it in a solution made by dissolving ½ teaspoonful of sodium hydroxide ($NaOH$), that is, caustic soda, in a beaker half full of boiling water (H_2O).

This hot solution will dissolve all the grease and oxide and leave the coin perfectly clean. This done, put the coin into a smaller beaker and pour over it the nickel solution, which you have allowed to get cold. Finally, put a strip of clean zinc (Zn) into the solution and let it stand for 5 or 10 minutes. The nickel (Ni) will be deposited upon the coin and the latter will be *nickel-plated.*

Tin, the Soft, Malleable Metal. "Tin" is an old Anglo-Saxon word, and *stannum* is its Latin name, so from this we get the symbol for it (Sn). This metal was one of the earliest known, as the fact that it has been found in Egyptian tombs goes to show. In the days of the early Greeks tin (Sn) was found only in the British Isles and, hence, these were called the *Tin Islands.* It is a soft, white metal and so malleable that it can be rolled into exceedingly thin sheets which we call *tin-foil.*

The chief ore it is contained in is *cassiterite* (SnO_2), or *tin-stone*, as it is commonly called, and this is found in the Straits Settlement, Bolivia, England, and Nigeria. The metal is extracted from its ore by pulverizing the latter, washing it, and then roasting it to drive out the impurities, after which it is smelted with coke (C). The chemical compounds formed of tin (Sn) include the chlorides, the oxides, stannic acid (H_2SnO_3), etc.

An Experiment with Tin. This is an experiment in which stannous chloride ($SnCl_2$) is used to separate the metallic mercury (Hg) that is in a mercuric chloride ($HgCl$) solution. Make a little stannous chloride ($SnCl_2$) solution by dissolving ½ teaspoonful of granulated tin (Sn) in a test tube half full of hydrochloric acid (HCl). Next dissolve ½ teaspoonful of mercuric chloride ($HgCl$), the common name of which is *corrosive sublimate*, and which you want to handle very carefully, as it is *poisonous*, in a test tube half full of water (H_2O). Now add a few drops of the tin solution, and the metallic mercury (Hg) will appear as a grey powder and fall to the bottom of the test tube.

Lead, the Heavy Metal. Lead (Pb) like tin (Sn) is one of the oldest known metals. It gets its symbol from *plumbum*, which is its Latin name, and it was widely used by the ancient Romans for weights, utensils, and water pipes. It is found free in small quantities, but the greater part of it is extracted from a mineral called *galena* (PbS), which is *lead sulphide* (PbS). To get rid of the sulphur (S), the ore is roasted, and this drives it off. As lead (Pb) does not rust, it is largely used for plumbing, and because it does not react with hydrochloric acid (HCl) or dilute sulphuric acid (H_2SO_4), it is used for making vessels for holding these acids. It is also the metal that is used for making storage battery plates.

In combination with other elements it forms lead oxides, one of which is *minium or red lead* (Pb_3O_4), and lead carbonate ($PbCO_3$), which is *white lead*, both of which are used for making paint; lead nitrate ($Pb(NO_3)_2$)[1], lead acetate ($Pb(CO_2CH_3)_2H_2O$), that is, *sugar of lead*, so-called because it has a sweet taste, lead sulphate ($PbSO_4$), and lead sulphide (PbS).

How to Make a Lead-Tree. You can make the pretty vegetable-like growth called a *lead-tree*, or Arbor Saturni, by precipitating the lead (Pb) from one of its salts by means of zinc (Zn). Dissolve 1 ounce of powdered lead nitrate ($Pb(NO_3)_2$) in a pint bottle of water (H_2O). Now tie a bit of granulated zinc (Zn) to one end of a thread, then fix the other end to a cork and suspend it in the lead solution

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[1] This salt is made by treating lead (Pb) with nitric acid (HNO_3).

so that it will be in the center of the jar. In the course of several hours the lead (*Pb*) will be slowly deposited in the form of a tree, as shown in Fig. 124. While this action is going on, the zinc (*Zn*) will pass into the solution and so exchanges places with the lead, thus:

$$\begin{array}{ccccccc} Zn & + & Pb(NO_3)_2 & = & Zn(NO_3)_2 & + & Pb \quad \downarrow \\ \text{Zinc} & & \text{Lead nitrate} & & \text{Zinc nitrate} & & \text{Lead} \end{array}$$

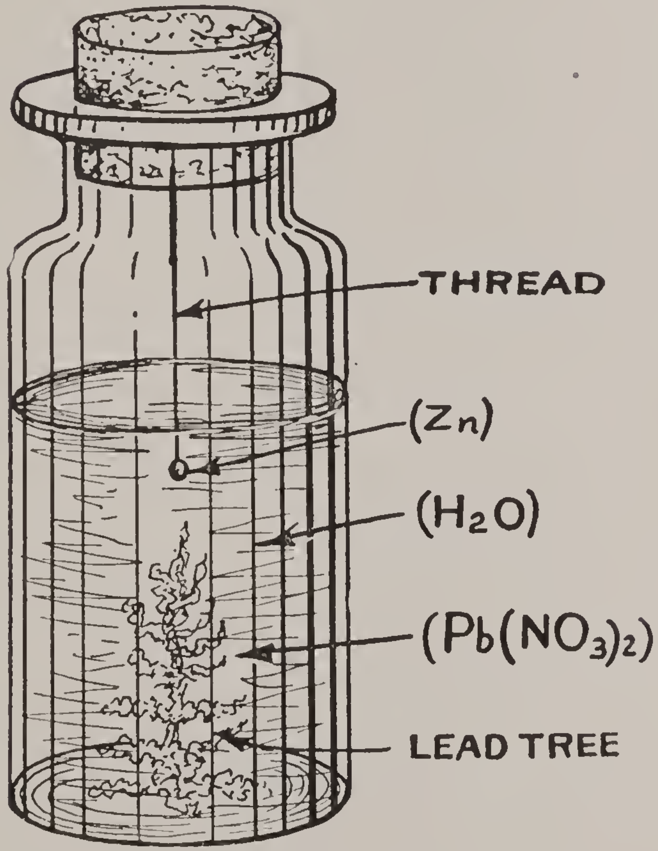

FIG. 124.—How to Make a Lead-Tree.

Copper, the Prehistoric Metal. After the *stone age* came the *copper age*, and the reason that copper (*Cu*) was the first metal to be used by the pre-historic races is because it is found free in nature, and, what was also fortunate for them, it was soft enough to be welded into shape while it was cold. In the early Roman days it was brought from an island in the Mediterranean called *Cyprus*, and they named this metal *Cyprium aes*, which means *Cyprium brass.* Then as time went on it was called *cuprium*, then it degenerated into *cuper*, and we call it by the good old Anglo-Saxon name of copper (*Cu*). It is not only found free in considerable quantities but it occurs plentifully in many kinds of ores.

It is, as you probably know, a tough, reddish metal, and as it does not rust to any great extent and wears well it is widely used by all nations for making coins of the smallest value. For the same reason, it is also used for making shells for rifles and guns, for cooking utensils, and as it is the next-best conductor of electricity,[1] it is especially useful for electric wires and apparatus.

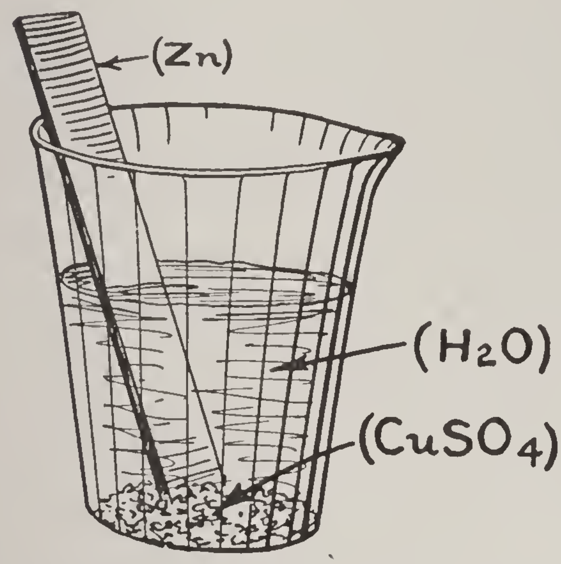

FIG. 125. How to Electroplate with Copper.

With other elements it forms various compounds, and these belong to two distinct series, the first of which is called *cuprous compounds* and the second *cupric compounds.* These compounds include the chlorides, bromides, oxides, hydroxides, carbonates, cyanides, acetates, sulphates, and sulphides. Cupric hydroxide ($Cu(OH)_2$) is used with ammonium hydroxide (NH_4OH) to form a compound that has this formula ($Cu(NH_3)_4(OH)_2$). Cellulose ($C_6H_{10}O_5$) in the form of paper or cotton will dissolve in this solution, and when this is forced through minute holes in steel plates, threads of *artificial silk* are formed.

An Experiment with Copper. Put 1 ounce of cupric sulphate ($CuSO_4$), or *copper sulphate, blue vitriol,* or *bluestone,* as it is variously called, in a beaker of warm water (H_2O), stir it well, and set a strip of zinc (Zn) in it, as shown in Fig. 125. Very soon the zinc (Zn) will be plated with copper

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[1] Silver (Ag) is the best conductor.

(Cu), and some of the zinc (Zn) will take the place of the copper (Cu) in the solution, and it becomes zinc sulphate ($ZnSO_4$) thus:

$$\underset{\text{Zinc}}{Zn} + \underset{\text{Copper sulphate}}{CuSO_4} = \underset{\text{Zinc sulphate}}{ZnSO_4} + \underset{\text{Copper}}{Cu}$$

Bismuth, the Easily Fusible Metal. Although bismuth (Bi) is found free in nature, still it is not a commonly known metal. Just why it is called bismuth (Bi) seems not to be known, and Agricola, who discovered it in 1529 called it *wiessmatte*, which means a *blooming meadow*, because of the variegated colors it shows when it is tarnished. It is found in ores formed of bismuth trioxide (Bi_2O_3), and bismuth trisulphide (Bi_2S_3), or *bismuth glance* as it is called. Bismuth (Bi) is a pinkish-colored metal, very brittle, melts at a low temperature, and has the peculiar property of expanding when it cools.

Bismuth (Bi) does not tarnish when exposed to air, and when heated to redness it burns and forms bismuth trioxide (Bi_2O_3). It united with fluorine (F), chlorine (Cl), bromine (Br), iodine (I), and nitrogen (N). Bismuth nitrate ($Bi(NO_3)_3, H_2O$) is the best known and most important salt of this metal, and this forms the well-known cosmetic so largely sold for beautifying ladies' complexions under the name of *pearl white*.

Experiments with Bismuth. Drop some finely powdered bismuth (Bi) into a jar of chlorine (Cl) and it will take fire.

Put a bit of bismuth (Bi) on a piece of charcoal (C) and heat it. It will then form a yellow film, which is bismuth trioxide (Bi_2O_3).

Antimony, the Metal that Expands. This metal was discovered by Valentino in the latter part of the 15th century and it gets its name from two Greek words which mean *against* and a *monk*, because some monks were poisoned by medicine made from it. Antimony (Sb) is found free in nature, but the chief supply comes from an ore called *stibnite* — hence the symbol (Sb),— which is black antimony sulphide (Sb_2S_3); and this ore is roasted in the air to drive out the sulphur (S).

Antimony (Sb) has a silver-white color and, like bismuth (Bi), it is brittle, melts at a low temperature, and expands on cooling. For the last reason, it is mixed with lead (Pb), which contracts, for making type, and very sharp edges result. The well-known remedy called *tartar emetic* is a chemical compound formed of potassium (P), antimony trioxide (Sb_2O_3), tartaric acid ($C_4H_6O_6$), and water (H_2O).

There are quite a number of antimony compounds, and these include stibine (SbH_3), which is *antimoniuretted hydrogen*, made by the action of zinc (Zn) and hydrochloric acid (HCl) on some compound of antimony (Sb) which is soluble. Then there are the halides, oxides, and the sulphides, antimony salts, such as antimony nitrate ($SbNO_3$) and antimony sulphate ($Sb(SO_4)_3$), antimonic acid (H_3SbO_4), and the sulphantimonites and sulphantimonates.

Experiments with Antimony. Heat a bit of antimony (Sb) by laying it on a piece of charcoal (C), and bringing the flame of your alcohol lamp or Bunsen burner to bear on it with a blowpipe, as shown in Fig. 126, and the melted bead of metal will show a network formed of antimony trioxide (Sb_2O_3). Now melt a small amount of antimony (Sb) in a crucible, grip the crucible with a pair of tongs and, holding it out at arm's length, pour the antimony upon the ground. It will form smoking globules that rebound and rush up like lava thrown out of a volcano.

Mercury, the Liquid Metal. This strange metal which is a liquid at ordinary temperatures was a great favorite of the old alchemists, for other metals except iron (*Fe*) and platinum (*Pt*) will dissolve in it, and in this way *amalgams* are formed. The Latin name for mercury (*Hg*) is *hydrargyram*, and this is the source of its symbol (*Hg*).

FIG. 126.—How to Heat Antimony with a Blow-Pipe.

It is a bright, silvery-white metal, and it is found both free in nature, as little drops, and in combination with sulphur (S) in mercuric sulphide (HgS), or *cinnabar*, as it is called. To get the mercury (Hg) out of the latter, the ore is roasted, causing the oxygen (O) of the air to combine with the sulphur (S) in it. Sulphur dioxide (SO_2) is formed, and the mercury (Hg), which passes off as a vapor, is caught, and then condensed again, thus:

$$\underset{\substack{\text{Mercuric}\\ \text{sulphide}}}{HgS} + \underset{\text{Oxygen}}{O} = \underset{\substack{\text{Metallic}\\ \text{mercury}}}{Hg} + \underset{\text{Sulphur dioxide}}{SO_2}$$

Of the compounds of mercury (Hg), the most common ones are mercuric oxide (HgO); mercurous chloride ($HgCl$), or *calomel*, as it is called in medicine; mercuric chloride ($HgCl_2$), which is corrosive sublimate; mercurous iodide (HgI), mercuric iodide(HgI_2), and mercuric sulphide (HgS), which brings us back to the ore we started from.

An Experiment with Mercury. Clean a small rod of zinc (Zn) with a little dilute sulphuric acid (H_2SO_4), and then roll it around in a few drops of mercury (Hg), and it will be coated all over with the latter, or amalgamated, as it is called. Zinc battery plates are *amalgamated*, and the layer of mercury (Hg) prevents a *local action* from being set up between the impure particles in the zinc (Zn) and the atoms of the latter, which lessens the current output of the battery.

Silver, the Queen of Metals. The Latin name for silver (Ag) is *argentum*, and since S is the symbol for sulphur, Ag is used as the symbol for silver. It is the least valuable of the precious metals, but owing to its many good properties it is widely used in the arts. It is found free in nature and also in combination with sulphur (S) as silver sulphide (AgS), and this, in turn, is often found in galenite (PbS). It is extracted from these ores by Parke's process.[1]

Silver (Ag) does not tarnish in air, but the fumes of sulphur (S) coat it with a thin film. It is the best of all known conductors of electricity, but as it is much more costly than copper (Cu), and the latter is nearly as good, it is not used for this purpose. Finally, it is the favorite metal for coinage, and much of it is made into silverware. It is also largely used for electroplating and in photography.

The chief compounds that are formed with it are silver chloride ($AgCl$), silver nitrate ($AgNO_3$), silver bromide ($AgBr$), and silver iodide (AgI). There are other salts of silver (Ag) and many complex compounds of it, but these need not be gone into here.

An Experiment with Silver. Since silver (Ag) is too soft to be used alone for coins, it is mixed, or alloyed, as it is called with a little copper (Cu). Hold a silver dime in a gas flame with your pair of forceps until it is quite hot, then put a couple of drops of water (H_2O) on it and let it cool, and there will be a black spot on it. This spot is cupric oxide (CuO), that is, *copper oxide*, and it is caused by the copper (Cu) in the coin combining with the oxygen (O) of the water (H_2O).

Gold, the King of Metals. Gold (Au), that wonderful yellow metal, was the warp, and the struggles of those who sought it were the woof of which some of the most thrilling scenes in the world's history were woven. Long ages ago

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[1] A description of this process will be found in Alex. Smith's "Inorganic Chemistry," published by the Century Co., New York.

it was called by the Latin name of *aurum*, and it is from this that we get the symbol, Au.

Gold (Au) is usually found free in nature — very often in quartz sand and also encased in quartz. It is separated from the former by washing and from the latter by mercury (Hg), which forms an amalgam with it; when this is gently heated, the mercury (Hg) passes off as a vapor and the gold (Au) remains behind. Gold (Au) is a soft metal and so malleable that it can be beaten into leaves of such exceeding thinness that it takes a quarter of a million of them to make a pile 1 inch high.

Gold (Au) will not set hydrogen (H) free from acids, it is not affected by air, and it will not dissolve in any kind of acid except *aqua regia*, a mixture of hydrochloric acid (HCl) and nitric acid (HNO_3). It has always been used for jewelry and for coinage, and to make it hard enough for these purposes it must be *alloyed* with a little copper (Cu). It is also used for making gold leaf, for gold plating, and formerly to a considerable extent in photography.

Gold (Au) combines directly with chlorine (Cl), and when dissolved in aqua regia it gives chlorauric acid ($HAuCl_4.4H_2O$). Auric chloride ($AuCl_3$), or *gold chloride*, which was once much used for toning photographic prints, is made by heating chlorauric acid ($HAuCl_4.4H_2O$); the latter gives up its hydrogen chloride (HCl) and leaves the red crystals of auric chloride ($AuCl_3$) behind. Gold (Au) also combines with bromine (Br), and there are several other compounds of it.

An Experiment with Gold. Get a sheet of gold leaf of a sign-painter, or a painter's supply house, and hold it up before a white light. You will observe that the light passes through the gold leaf quite easily.

Platinum, the Regal Metal. The word platinum (*Pt*) comes from the Spanish *platina*, which means silver (*Ag*), and since these metals bear a decided resemblance to each other in the matter of color, it is quite likely that the regal metal was taken for the queen of metals when it was first found. Platinum (*Pt*) is chiefly found free in the gravel of river beds, and most of it comes from the Ural Mountains. It is quite hard, cannot be melted in the flame of a Bunsen burner, but does so in an oxyhydrogen flame or the electric arc.

Platinum (*Pt*) is a very heavy metal, a piece of it weighing nearly three times as much as a piece of iron (*Fe*), and twice as much as a piece of lead (*Pb*) of the same size. It has a very small chemical activity, by which is meant that it resists the action of most substances, and hence it is largely used in making crucibles, evaporating-dishes, and other apparatus required by the chemist.

One of the chief compounds made of platinum (*Pt*) is platinic chloride ($PtCl_4$), and this is done by dissolving the metal in *aqua regia* and then evaporating the solution, causing the acids to pass off.

How Alloys are Made. The name alloy, is given to two or more different metals when they are melted together. While an alloy thus produced is simply a *mixture* of the metals, it possesses properties entirely different from either or all of them.

Alloys of Magnesium and Aluminum. Aluminum (*Al*) is a metal that is hard to work in a lathe and also to polish, for the reason that it sticks to the tools, but when 2 per cent of magnesium (*Mg*) is mixed with 98 per cent of it, they form an alloy called *magnalium* that is free from these difficulties.

Alloys of Iron and Steel. Pure iron (*Fe*) is quite soft and malleable, but what we call *cast iron* is really an alloy which contains a small per cent of carbon (*C*), sulphur (*S*), silicon (*Si*), and other substances, and it is these that make it grey and brittle. Wrought iron is purer, and for this reason it is malleable. Steel is simply iron (*Fe*) that has more carbon (*C*) in it than cast iron, and when it is heated to a certain temperature and allowed to cool slowly, it gets very hard and springy.

By mixing in various other metals, a wonderful series of *steel alloys* are produced; thus when 7 to 20 per cent of manganese (*Mn*) is added to steel it is called *manganese steel.* This alloy is exceedingly hard, and so it is used for making burglar-proof safes, the jaws of stone crushers, and railway frogs. When 1 per cent of chromium (*Cr*) and 15 per cent of vanadium (*V*) are mixed with steel, it makes an alloy called *chrome-vanadium steel*, and this not only has great tensile strength but it will bend double before it will break. It is used for the connecting-rods of engines and the frames and axles of motor-cars.

An alloy with 20 per cent of tungsten (*W*), ¾ per cent carbon (*C*), 3½ per cent of chromium (*C*r) and 1½ per cent of vanadium (*V*), and the rest steel, makes an alloy called *high-speed steel.* This is used to make tools for lathes and other machines, and these will cut steel very fast and without losing their temper even when the friction heats them red-hot. When 2 to 4 per cent of nickel (Ni) is added to steel it makes an alloy called *nickel-steel*, and this alloy is not only hard and springy but one that sea-water (H_2O) has but little effect upon. For these reasons it is used for propeller shafts, ships' cables that must be placed underseas, armor-plate for battle-ships, etc.

Another nickel-steel alloy is called *invar steel*, and this is made of 36 per cent of nickel (Ni) and the rest steel. While all other metals and alloys will either expand or contract on cooling, invar steel remains of the same dimensions under practically all degrees of temperature. Hence it is largely used for clock pendulums and measuring scales, the lengths of which must remain unchanged under all conditions.

Alloys of Tin and Lead. *Common solder* for roofing purposes is made of 50 per cent of tin (Sn) and 50 per cent of lead (Pb). *Fine solder* for soldering tinware is formed of 33⅓ per cent of lead (Pb) and 66⅔ per cent of tin (Sn). *Type metal* is made of 25 per cent of antimony (Sb) and 75 per cent of lead (Pb). *Pewter* is an alloy the components of which consist of 20 per cent of lead (Pb) and 80 per cent of tin (Sn). An alloy that has a ½ of 1 per cent of arsenic (As) in it and 99½ per cent of lead (Pb) is used for making shot and bullets. *Wood's metal* is an alloy that melts at a much lower temperature than that at which water (H_2O) boils, and it is used for electric fuses and safety plugs for steam boilers and automatic sprinklers. It is made of 1 part of tin (Sn), 1 part of cadmium (Cd), 2 parts of lead (Pb), and 4 parts of bismuth (Bi).

Alloys of Copper. One of the earliest alloys to be used is bronze, and it is made of varying proportions of lead

(Pb), tin (Sn), zinc (Zn) and copper (Cu), depending on what it is to be used for. *Brass*, of which there are also several varieties, is made of from 18 to 40 per cent of zinc (Zn) and the rest copper (Cu). *German silver* has no silver (Ag) in its make-up but is formed of 20 per cent of zinc (Zn), 20 per cent of nickel (Ni), and 60 per cent of copper (Cu); it looks something like silver (Ag) and does not easily tarnish; at one time it was much used for making spoons.

Gun metal is an alloy that consists of 10 per cent of tin (Sn) and 90 per cent of copper (Cu); these proportions give a metal that has a rich brownish-black color and is much used for making art objects and the like. *Monel metal* is made of 50 per cent of nickel (Ni) and 50 per cent of copper (Cu); it is largely used for sheet metal work. Finally, an alloy formed of 30 per cent of manganese (Mn) and 70 per cent of copper (Cu) makes a hard bronze that is used for the propellers of ships.

Silver Alloys. The alloy of silver (Ag) that is used for coins is 10 per cent of copper (Cu) and 90 per cent of silver (Ag), while that used for silverware is made of 20 per cent of copper (Cu) and 80 per cent of silver (Ag). In both cases the copper (Cu) is put in to make the metal harder.

Gold Alloys. To make gold (Au) hard enough for coinage and goldware, it is also alloyed with copper, and the proportion of gold used is measured by a unit called a *carat.* The carat in general use is equal to 3.168 grains, or 205 milligrams. Pure gold (Au) is said to be 24 *carats fine.* British sovereigns are 22 *carats fine*, and hence have 2/24 or 1/12 of copper (Cu) in them. American and French gold coins are 21 6/10 carats fine, that is, they contain 10 per cent of copper (Cu) and 90 per cent of gold (Au).

How Amalgams Are Made. An *amalgam* is an alloy in which one metal is dissolved in another metal. The word amalgam comes from two Greek words that mean *soft mass.* Mercury (Hg) is used to dissolve the other metal, and all the common metals will dissolve in mercury (Hg), although platinum (Pt) and iron (Fe) do so to the least extent.

A Sodium Amalgam. When 1 per cent of sodium (Na) is dissolved in 99 per cent of mercury (Hg), an amalgam is formed that is a soft mass, but when 2 per cent of sodium (Na) is dissolved in 98 per cent of mercury (Hg) it forms a solid mass. When metallic sodium (Na) is to be used, it is often better to make an amalgam, as the mercury (Hg) will not affect its action and it can be handled easier.

Zinc Amalgam. When the zinc (Zn) plates used in batteries are rubbed with mercury (Hg), the molecules of the former that are on the surface are mixed with those of the latter, and the plates are then said to be amalgamated. The amalgam thus formed prevents local currents from being set up between the impure particles in the zinc (Zn), and the zinc (Zn) itself, and also keeps the acid solution from eating the zinc (Zn) away so rapidly.

Tin and Zinc Amalgams. Formerly mirrors were made by coating glass (Na_2O,CaO,SiO_2) with a tin amalgam formed of 1 part of tin (Sn), 1 part of lead (Pb), and 2 parts of bismuth (Bi) dissolved in 4 parts of mercury (Hg). Mirrors are now made by coating them with a silver nitrate ($AgNO_3$) solution.

When 1 part of tin (Sn) and 2 parts of silver (Ag), or gold (Au), are dissolved in 2 parts of mercury (Hg), an amalgam is formed of these metals that will harden and expand. For these reasons, this amalgam is used by dentists for filling cavities in teeth.

A cheap amalgam for filling teeth is made by dissolving 1 part of pulverized zinc (Zn) in 2 parts of mercury (Hg).

CHAPTER X.

CHEMISTRY SIMPLY EXPLAINED

QUITE the least-interesting part of chemistry to the beginner is the theory of it, by which is meant just how and why different substances act and react on each other and produce other substances of an entirely different nature.

Now while making experiments of any kind, but especially those in chemistry, is the most fascinating part of the work, and you can get the same results if you follow the instructions given whether you know anything about the principles that underlie them or not, still, to work in this way is to miss much of the fun and more of the interest in it. So in this chapter I shall tell you about some things in chemistry which you cannot see with your physical eye but which you can visualize with your mind's eye.

What Matter Is. First of all, as you probably know, the world and everything in, on, or around it is made up of minute particles of different substances, and it is these that form the material bodies which we call *matter*. Now matter of whatever kind has certain *properties*, and the first two of these are called by the long names of *indestructibility* and *impenetrability*; then solids have half a dozen other properties, and these are *hardness*, *malleability*, *ductility*, *brittleness*, *elasticity*, and *flexibility*. Fluids include liquids and gases; the former are called *inelastic*, while gases have *elasticity*.

What the Properties of Matter Are. *Indestructibility* is a word used to mean that matter cannot be created or destroyed. It is easy to change the shape of solid bodies, to make two or more substances into an entirely different substance, to make a solid into a liquid, and the latter into a gas; then, the other way about, gases can be liquefied and these in turn can be solidified, though these latter processes are much more difficult than the former. But in any event there is exactly the same amount of matter left as there was before the operation.

Impenetrability means simply that no two particles of matter can be in the same place at the same time. This is perfectly obvious with solid bodies but not so much so with liquids and with gases. Now here are two experiments, the first of which shows easily enough the truth of the first statement, and the second *seems* to show the fallacy of it.

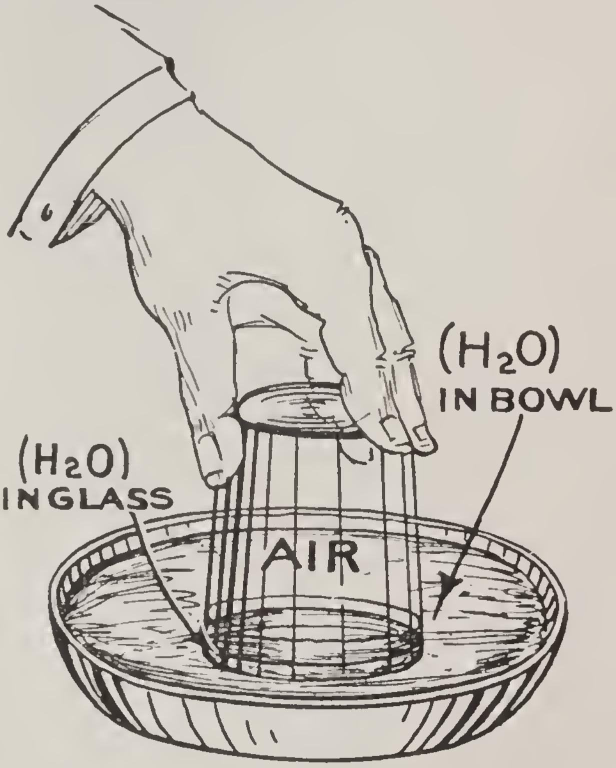

FIG. 127.—Clearly Two Bodies Cannot Occupy the same Space at the Same Time.

First Experiment. Hold a tumbler, or a beaker, by the bottom and put the open end into a soup-plate half full of water (H_2O), as shown in Fig. 127. You will see now that the level of the water (H_2O) in the glass is very much lower than it is outside of it. It is evident that the air in the glass prevents the water (H_2O) from rising in the glass and taking up the same space that the air occupies.

Second Experiment. Now here is an experiment of a different kind. Fill a beaker, or a tumbler, brimful of water (H_2O) and then take a tablespoonful of pulverized sugar ($C_{12}H_{22}O_{11}$) and slowly let it fall into the water (H_2O), a particle at a time, as shown in Fig. 128, so that the bubbles of air which stick to them will rise to the surface, and also to give the sugar ($C_{12}H_{22}O_{11}$) sufficient time to dissolve.

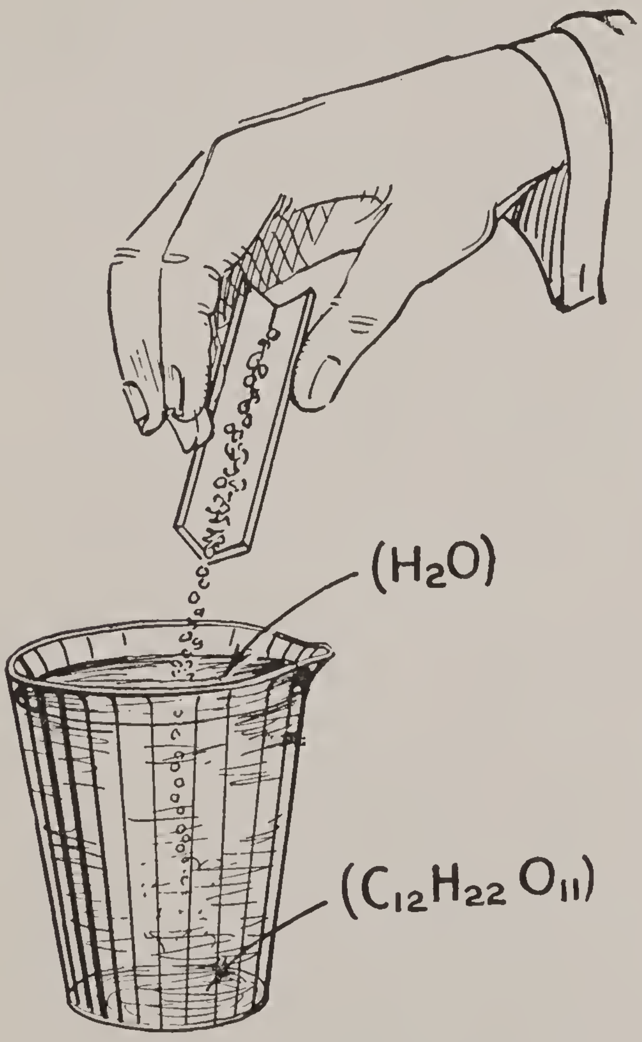

FIG. 128.—This Experiment seems to Show that Two Bodies Can Occupy the Same Space at the Same Time.

Curiously enough, as you put the sugar ($C_{12}H_{22}O_{11}$) into the water (H_2O), the latter does not increase in volume, and so run over the top of the glass. Now while it is true that two bodies cannot occupy the same space at the same time, these two substances seem to do this very thing. The explanation, however, is that the particles, or *molecules*, as they are called, of which the water (H_2O) and the sugar ($C_{12}H_{22}O_{11}$) are formed, are widely separated, and when the latter is dissolved in the former, the spaces between the particles of the one are filled up by the particles of the other.

Hardness is that property of a body which makes it tend to resist any change in shape. The degree of hardness is found by scratching one substance with another; thus iron (Fe) will scratch lead (Pb); glass (Na_2O, CaO, SiO_2) will scratch iron (Fe), and the diamond (C) will scratch glass (Na_2O,CaO, SiO_2).

Malleability is that property of matter which permits a solid, such as iron (Fe), silver (Ag), gold (Au), or platinum (Pt) to be rolled out into sheets. When red-hot, iron (Fe) is very malleable. This is also true of steel, which is iron (Fe) with a small amount of carbon (C) in it. In this heated state, steel is rolled into sheets, rails, girders, etc. You have seen in the foregoing chapter that gold (Au) possesses this property to such an extent that it can be beaten into sheets so thin it will take 250,000 of them to make a pile 1 inch high.

Ductility is a first cousin to malleability, for it is that property which permits a metal to be drawn out into a slender thread, or *wire*, as it is called. When metals of various kinds are drawn into wire, they have a greater strength than they would have if a cross-section of the same size were cut from a strip. Evidently the molecules of which they are formed are rolled closer together. *Brittleness* is just the opposite of malleability, and it is this property of matter that makes a sheet of glass crack when it is struck, and some metals are so brittle that they break when you try to bend them.

Elasticity is that property of solids and gases which makes them return to their original shape when they have been sprung by the application of a force which is then released, as for instance a steel spring when it is bent or twisted, or a gas when it is compressed. *Flexibility* is that property which enables a body to be bent out of shape, without breaking.

The Three Common Forms of Matter. The three forms of matter that we are best acquainted with are the solid, liquid, and gaseous, and, as I explained before, these may be changed one into the other. As an illustration, water (H_2O) at ordinary temperatures is a liquid, but when it is heated to a high-enough temperature it becomes a gas, and when it is cooled to a low-enough temperature it becomes a solid. Since metals can be liquefied, and these liquids converted into gases, and gases can be liquefied and then changed into solids, it is safe to conclude that all the elements and some of the compounds follow the same laws.

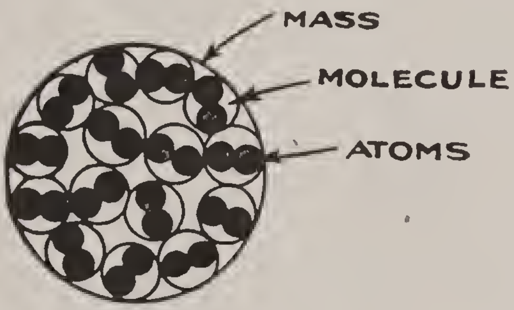

FIG. 129.—How Atoms Form the Molecule, and Molecules the Mass.

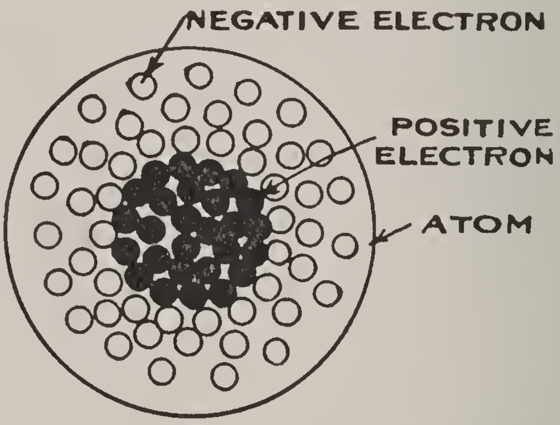

FIG. 130.—How the Negative and Positive Particles of Electricity form the Atom.

What Matter Is Built Up of. The way we always see matter, even with the highest-power microscope, is in the form of a *mass*. A mass of any kind is built up of *molecules*, which in turn are built up of *atoms*, and, finally, the latter are built up of *electrons*. The way in which atoms form molecules and the latter form masses is shown in the diagram in Fig. 129, and the way an atom is built up of electrons is shown in Fig. 130.

The center of an atom is built up of a number of compact *positive* charges of electricity called *electrons*, which are represented by the black dots,and these are surrounded by less compact charges of *negative* electricity, also called *electrons*, which are shown by the white circles. It is the number of positive charges of electricity an atom has that determines whether it is iron (*Fe*), or sulphur (*S*), or oxygen (*O*), etc.

Now the main things for you to remember in chemistry, until you get into that part of it in which electric currents are used, is that an atom is the smallest particle into which matter can be divided without changing its nature, and two or more atoms bound together by chemical affinity form a molecule.

The molecules of a solid are held together by a much stronger attractive force than they are in a liquid and while they are constantly vibrating, that is, moving to and fro, still they cannot move freely about, and so the body keeps its shape. In a liquid they move about quite freely, but still they are held together by an attractive force. It is, however, the force of gravity that pulls them down, and so makes the liquid take on the shape of the vessel that holds it.

The action of a gas is quite different, for all the molecules repel each other exactly as if they were charged with the same sign[1] of electricity, and, hence, they shoot out in every direction. This is the reason why a gas when it is compressed exerts a force against all of the inside surface of a tank, or a gas-bag, that holds it.

What the Elements Are. An *element* is a mass of matter that is built up of molecules, which are, in turn, formed of atoms all of which are of a single kind. Two or more of these atoms may be linked together to make up a molecule, as shown in Fig. 131; this is the way that the molecules which form oxygen (O) are linked together. Sometimes the element itself is changed when an extra atom of the same kind is added to those that form the original molecule; thus if 3 atoms of oxygen (O) are linked together, as shown in Fig. 132, then the molecule becomes ozone (O_3).

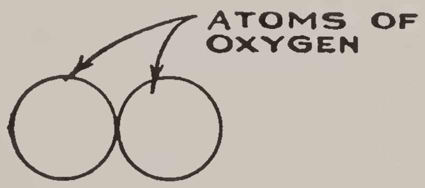

FIG. 131.—Two Atoms of Oxygen Make a Molecule of Oxygen.

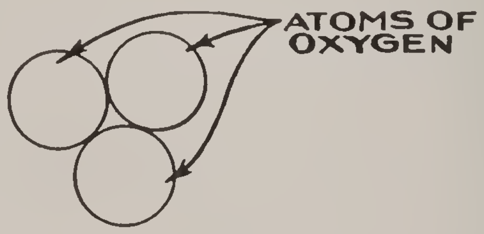

FIG. 132.—Three Atoms of Oxygen Make a Molecule of Ozone.

Now there are 83 different elements known at the present time, and a table of these will be found at the back of this book, so that you can conveniently turn to it when you want to find the *symbol* of one of them.

How the Elements Got Their Names. The names of the elements are interesting, but it is not at all easy to trace

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[1] That is, either with + electricity or — electricity, for like signs repel each other.

the older ones. Thus iron (*Fe*) and gold (*Au*) are so old that their names are shrouded in obscruity. Many of the elements that have been discovered in recent times get their names from various Greek and Latin words.

Thus chlorine (*Cl*) comes from the Greek word which means *yellowish-green*, as this is the color of the gas. Bromine (*Br*) gets its name from the Greek word which means *stench*, and it is very aptly named, too, for it is a very smelly gas. Hydrogen (*H*) comes from two Greek words which mean *water*, and *to produce*, for water (H_2O) is produced by burning hydrogen (*H*) in air. In the same way, nitrogen (*N*) gets its name from the Greek word which means *nitre*.[1]

The recently discovered gases, helium (*He*), argon (*A*), neon (*Ne*), krypton (*Kr*), and xenon (pronounced ze-non); (*Xe*) are all named from Greek words meaning respectively: the *sun*, *inactive*, *new*, *hidden*, and *stranger*. Finally, some of the other elements are named from their properties, such as radium (*Ra*), and some are named in honor of the countries the chemists were natives of who discovered them, as for instance, scandium (*Sc*), etc.

What the Symbols Mean. In order that the name of each element would not have to be written out every time it was used, or where several of them are used together, as in *equations*, the names of them have been abbreviated and only the first letter or two — usually the first two where more than one is needed — are used to indicate an element; thus *O* stands for oxygen, *H* for hydrogen, *C* for carbon, and so on.

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[1] The common name of potassium nitrate (KNO_3) is *nitre*.

Where there are two or more elements that begin with the same letter (there are 10 that start with *C*), the second or third letter of the name is also used as, for instance, *Cl* stands for chlorine, *Ca* for calcium, etc. Then again some of the symbols are formed of the first and second letters of the Latin names of the elements as *Fe* for iron, since *ferrum* means iron, *Cu* for copper, since *cuprum* means copper, etc.

What the Symbols Show. Now whenever you come to a symbol, which let us suppose is *H*, you instantly know that it stands for hydrogen; if it is *O*, you know that it means oxygen, and the same with all the other symbols. Further, when you see two symbols linked together, thus, HO, you know that it means a compound is formed of the elements hydrogen (*H*) and oxygen (*O*). When two or more symbols are so linked together to indicate a compound, they make what is called a *formula*.

Wherever you see a formula, you will almost invariably find a number marked in little figures after one or the other, or both, or all, of the symbols thus, H_2O. This number gives you a great deal of information in a very small space — in fact it is the chemists' system of short-hand — for it tells you at a glance that the compound contains 2 volumes of hydrogen (*H*) and 1 volume of oxygen (*O*). This compound is water (H_2O), and so wherever you see the formula you will know exactly what the substance is that the elements have made when combined. Where there is only 1 volume of an element used, as of oxygen (*O*) in the formula for water (H_2O), the figure 1 is not placed after it, but there is understood to be 1 volume, for if there were 2 or more, the number would be added to show it.

While the word *volume* has been used in the above explanation, an experiment given in Chapter IV, in which you analyze water (H_2O) by passing a current of electricity through it, shows that it is made up of 2 volumes of hydrogen (H) and 1 of oxygen (O). It also means that when 2 atoms of hydrogen (H) combine with 1 atom of oxygen (O) they produce a molecule of water (H_2O), and so it is with all other compounds.

What Equations are. When two or more elements are made to combine with each other and two or more other elements or compounds are produced by the reaction, it is called an *equation*, because the quantities you start with and those that you get in the end are exactly equal. Take, for instance, the first equation I have given in this book in Chapter V under the caption of "How to Make Hydrogen without an Acid," which is

$$Zn + KOH = K + ZnO + H \uparrow$$

In this experiment, zinc (Zn) which is an element, is added to potassium hydroxide(KOH),that is, caustic potash, and which is a compound made up, as its formula shows, of potassium (K), oxygen (O), and hydrogen (H). Now when these react on each other, the oxygen (O) of the potassium hydroxide (KOH) combines with the zinc (Zn) and forms zinc oxide (ZnO), and this sets the hydrogen (H) free, and as this is a gas it passes into the air, so that the potassium (K) is left behind.

In the end, though, there is exactly as much zinc (Zn), potassium (K), oxygen (O), and hydrogen (H) as there was in the beginning, and to show that they are equal before and after the reaction, the equality sign is used. The equality sign, however, is not used by chemists nearly as much now as it formerly was in writing equations, a horizontal arrow having taken its place thus:

$$Zn + KOH \rightarrow K + ZnO + H \uparrow$$

If now instead of reading $Zn + KOH$ *equals* $K + ZnO + H$, you will read it $Zn + KOH$ *makes* $K + ZnO + H$, it will be just as sensible, though not quite so definite. As I have mentioned in the experiments that have gone before, where you find an *arrow* pointing *up* after a symbol in an equation, it means that the element or substance which has been set free is a *gas*, and, oppositely, where you find an arrow pointing *down*, it means that the element or substance which has been set free is a *precipitate.*

CHAPTER XI.

FIRE, FLAME, HEAT, AND LIGHT

WHILE, as Darwin has pointed out, man and monkey bear a very strong resemblance to each other, especially in their anatomical make-up, still they are as widely separated as the poles in their mental attributes. One of the most marked features which differentiate them in this latter respect is that the first knows how to make and to use *fire*, and the second shows an utter lack of any such knowledge.

That man began to use fire long before he could make it, there is not the slightest doubt, and he learned how to make it ages before that primitive race, called the Aryans appeared on the Iranian plateau, whence the early Hindus, Persians, Egyptians, and other races branched off. And it is curious to note that the Aryans used the word *agir* for *fire*, and that the Latin word for it is *ignis*, while we use the word *ignite* when we want to convey the meaning that we have lit, or started, a fire.

What Fire Is. When a substance combines slowly with oxygen (O), the process is called *oxidation*, and when it combines rapidly with oxygen (O), it is said to *burn*, and the process is called *burning*, or *combustion*, while the result of it in throwing off heat and light is called *fire.* The words fire, burning, and combustion, are, however, all generally used to mean the same thing, and that is that a chemical change is going on which produces both heat and light. Fire is, then, the chemical combination of a substance with some other substance that will support combustion, and for all ordinary purposes it is the air that supplies the oxygen (O) for the latter purpose, and this it does in unlimited quantities.

What Flame Is. When a solid substance burns that is formed chiefly of charcoal (C) or coke (C) or anthracite coal (C), the molecules of it are heated to incandescence, and while this gives out heat and a glowing light, it does not produce a flame. But when a substance that is formed of a gas or has gas in it burns in another gas that will support combustion they combine chemically, and the heated molecules flare up where the two gases come together, and this makes a flame, or blaze, as it is popularly called.

What Heat Is. A particle of matter just large enough to be seen by a microscope of fairly high power is formed of 8 or 10 billions of molecules. Now when a substance burns, the rapid chemical combination that takes place between it and the oxygen (O), or other substance which supports combustion, sets the molecules of which they are formed into violent vibration, that is, it gives them a rapid to and fro movement. In turn, these swiftly moving molecules strike those of the air, and when these reach the body they set the *thermal nerves* of the latter into vibration; these vibrations are transmitted to the brain and we get the sensation of *heat*. Or if they impinge on some inanimate mass of matter they make the molecules of it vibrate, and so it in turn gets hot.

What Light is. When a substance of any kind burns it gives out *light*. Now in the same way that the vibrating molecules of a burning substance set the air into motion they also set the *ether* into motion. The ether is a very thin and transparent kind of matter that fills all space which is not actually taken up by matter of other kinds, and it fills the pores of the densest metals. It is the substance by, in, and through which not only light, but all other electromagnetic waves travel. When the light waves reach the optic nerves of your eyes they set up the sensation of light in your brain, and they also have a very decided action on substances of various kinds, as you will see later in the chapter on photography.

Ways of Making Heat and Light. Our sun is, of course, the original source of all heat and light, and however these are produced, they are *directly* traceable to the sun. Heat is not always accompanied by light, but burning is; on the other hand, light may be had without any appreciable amount of heat accompanying it.

When heat is produced without light it means that the molecules of the substance that is heated are not vibrating fast enough to produce light. Oppositely disposed, the molecules of certain substances are capable of vibrating fast enough to set up light and yet not slow enough to produce heat, as, for instance, the phosphorescent light of a fire-fly, or a glow-worm, or the Geissler tube when it is energized by an induction coil.

The three chief ways of making heat are by friction, by chemical action, and by electricity, and in all cases light follows if the heat set up is sufficient to make the molecules of the substance vibrate fast enough. The only kind of heat and light that we are interested in now is that produced by chemical action.

How a Candle Burns. Light a candle and examine the flame of it through a piece of tinted glass and you will see that it consists of three parts, as shown in Fig. 133, and these are an inner dark part containing gas which is waiting its turn to be burned; a middle bright cone where the particles of carbon (*C*) are heated to incandescence and which gives the light; and a thin outer cone of blue flame which is in direct contact with the air and gives little or no light.

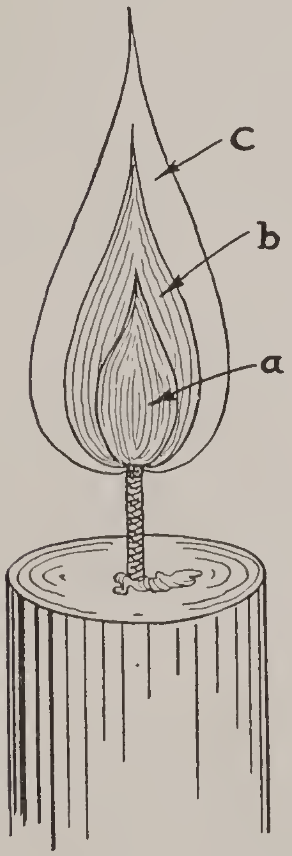

FIG. 133.—The Flame of a Candle.

Now this is what takes place when you light a candle. First, the heat melts the tallow, or wax, and this rises in the wick by *capillary attraction;*[1] as it reaches the tip of the wick where the heat is the greatest it is converted for the most part into a vapor, and this burns, which makes the flame. It is the carbon (*C*) of the tallow, or wax, which is raised to a white heat, while the hydrogen (*H*) of it burns with a blue flame outside of it and has no lighting power whatever.

How Ventilation Affects Combustion. Take a quart glass jar and fit a cork into the mouth of it; now bore two ¼-inch holes through the cork and push a glass tube 3 inches long through one of them and another tube 8 or 10 inches long through the other one so that the former will

[1] Any text-book on Physics will give you an explanation of this phenomenon.

extend just inside the jar and the latter will reach nearly to the bottom of it, as shown in Fig. 134.

Now put a lighted candle inside of the jar and you will observe that as it burns, vapor of water (H_2O) and carbon dioxide (CO_2) escape through the short tube. As they do so, fresh, cool air from the outside flows through the long tube into the jar and provides the necessary oxygen (O) to support the burning process. This experiment done, put your finger over the end of the long tube so that the fresh air is cut off, and the flame will soon begin to grow smaller, and finally it will go out altogether.

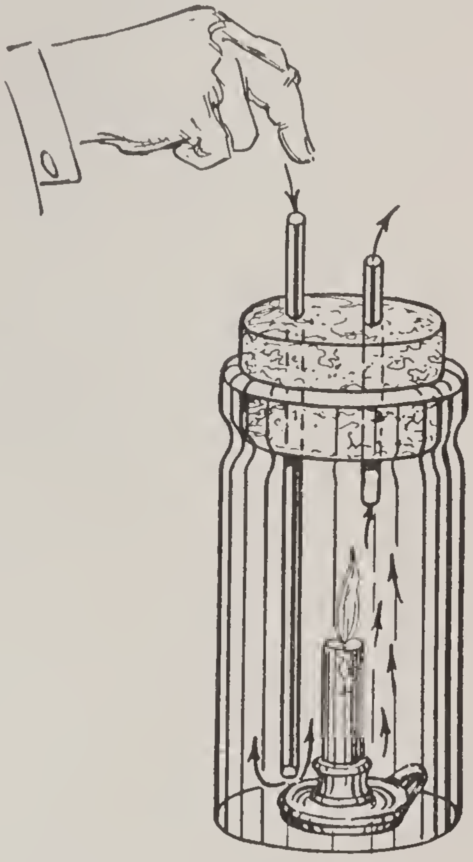

FIG. 134.—How Ventilation Affects Combustion.

How the Davy Safety-Lamp Works. Sir Humphrey Davy invented a safety-lamp, so that when miners who carried it entered a shaft where there was methane (CH_4), or *fire-damp*, as it is generally called, and which when mixed with air is very explosive, it would not ignite. His safety-lamp consists of a common oil lamp, the flame of which is surrounded by a wire gauze cover, as shown in Fig. 135. Now while enough oxygen (O) will reach the flame to keep it burning, the flame cannot get outside of the gauze to ignite the explosive gases.

To make an experiment which shows the principle of the safety-lamp, take a piece of fine wire gauze about 8 inches on the sides and which has about 10 meshes to the running inch and hold it in the flame of a candle with a pair of pliers, as shown in Fig. 136, and you will see that the flame remains beneath it. The reason it does not go through the gauze is because the wire of which it is made cools down the flame to such an extent that it puts it out, but the smoke and other gases of combustion pass through it readily enough.

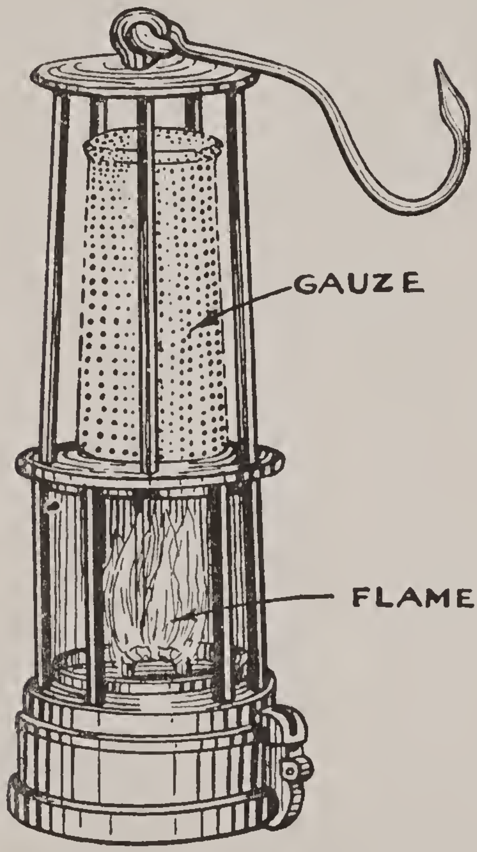

FIG. 135.—The Davy Safety-Lamp in Operation.

Now hold a lighted match above the gauze and the gases will ignite and will make another separate flame. The fact that a flame cannot pass through a piece of gauze is the principle, then, upon which the Davy safety-lamp is based.

How an Alcohol Lamp Burns. It often happens that a flame does not give out any useful light, and usually this kind of a flame is very hot. This is the case when pure hydrogen (*H*) burns in oxygen (*O*), generally obtained from the air. Now methyl alcohol (CH_4O) contains, as its formula shows, 4 times as much hydrogen (*H*) as it does carbon (*C*).

To produce a flame that has no useful lighting power, you must have not only a certain amount of oxygen (*O*) present, but it must be mixed with the hydrogen (*H*). When hydrogen (*H*) is mixed with oxygen (*O*), it gives a very hot flame, and this condition is fulfilled in a very simple manner in a lamp that burns alcohol (CH_4O).

How Oil and Gas Lamps Burn. When we want a bright light we must burn compounds that contain hydrogen (H) and carbon (C), and certain vegetable and mineral oils have them in the right proportions. Generally speaking, an oil to burn with a bright light must have a large amount of carbon (C) compared with the hydrogen (H) in it, and, further, its illuminating power also depends on the amount of oxygen (O) in which it burns, and this must be supplied gradually and from the outside.

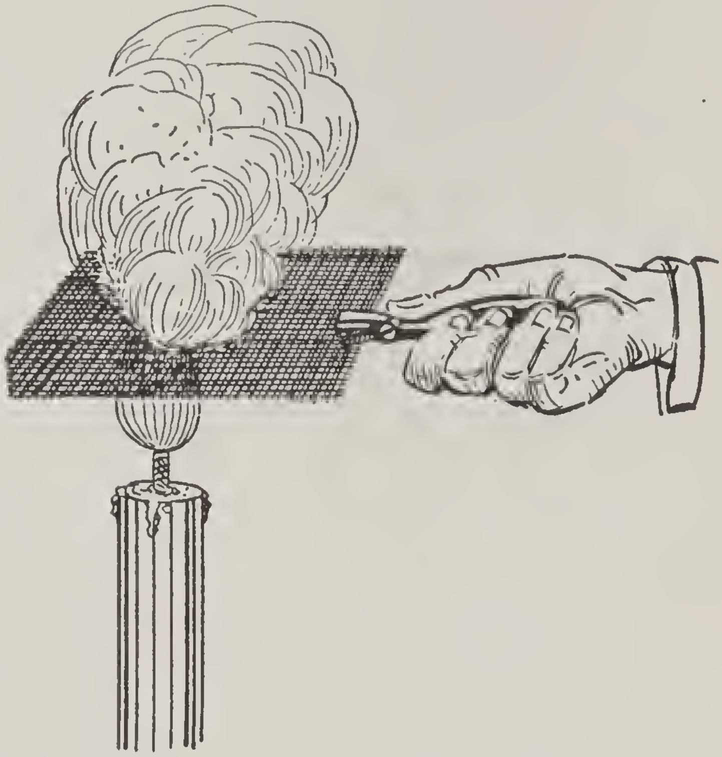

FIG. 136.—The Principle on Which the Davy Safety-Lamp Works.

Illuminating oils, such as kerosene ($C_{10}H_{22}C_{16}H_{34}$) and gases such as coal-gas (C_2H_4) do not contain oxygen (O), and these burn, therefore, with the brightest light. A tallow candle ($3C_{18}H_{35}, C_3H_5O$) contains a little oxygen

(O),and it burns with a less bright flame than oil or gas, while alcohol (CH_4O) contains considerable oxygen (O) in proportion to its hydrogen (H), and so it burns without any brightness whatever.

How a Bunsen Burner Works. While coal gas burns with a bright flame and with little heat in an ordinary burner, it can be made to burn with a flame that is hot, gives no light, and which does not smoke, by using what is called a *Bunsen burner*. It is so-called because it was invented by Bunsen, a German scientist who lived from 1791 to 1860.

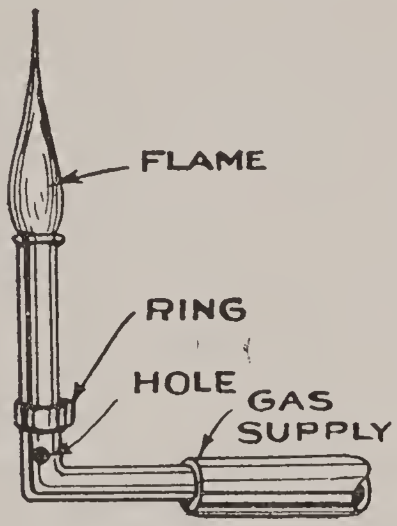

FIG. 137.—The Bunsen Burner.

A Bunsen burner in its simplest form consists of a brass tube ⅜ inch or ½ inch in diameter with a couple of air holes drilled through it near one end. A ring is slipped over the tube so that the amount of air which enters the holes can be controlled, and it is then connected with the pipe that supplies the coal gas, all of which is shown in Fig. 137. The purpose of the air-holes is to supply enough oxygen (O) to burn up the carbon (C) in the gas, and this makes the flame not only non-luminous but at the same time very much hotter.

EXPERIMENTS WITH A BUNSEN BURNER.

How to Light the Burner. The right way to light a Bunsen burner is to turn on the gas and then hold a match to one side of the top of it, upon which the gas will catch fire. If you hold the match over the top of the tube, the pressure of the gas may blow it out before the gas ignites.

The Luminous Flame of the Burner. Slip the ring over the holes in the burner so that it will cut off the air supply and, hence, the oxygen (*O*), and you will see that the flame gives out light, and if you hold a sheet of glass, or cardboard, over it, a film of carbon (*C*), which is ordinarily called *soot*, will be deposited on it.

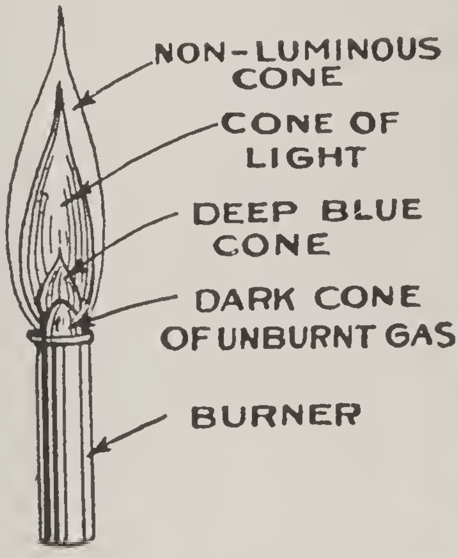

FIG. 138.—A Luminous Gas Flame.

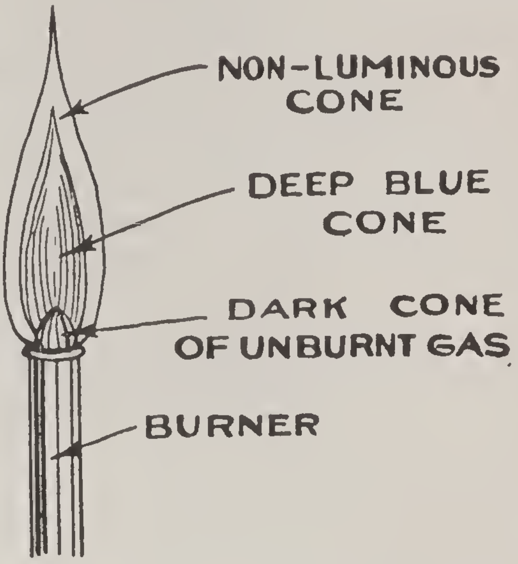

FIG. 139.—A Non-Luminous Gas Flame.

Now examine the flame and you will see that it is formed of four parts, as shown in Fig. 138; named, these are a dark blue cone next to the burner, next, a deep blue cone, then, a luminous cone containing glowing particles of carbon (*C*), and, finally, a colorless cone, or sheath, on the outside.

The Non-Luminous Flame of the Burner. After you have made the above experiment, slip the ring up and away from the air-holes, and as you do so you will see that the characteristics of the flame are changed. First of all, the glowing carbon (C) disappears and with it the light of the flame; at the same time the deep blue cone, which is the one that gives the heat, expands and takes its place, as shown in Fig. 139.

A closer examination of the flame will reveal the fact that the extreme tip is the hottest part of it, the next, or middle cone is not so hot, while the dark cone at the bottom is quite cool; this means that the gas is not burning at this point, and if you will put a small glass tube into it the gas will flow through it, and you can light it, as shown in Fig. 140.

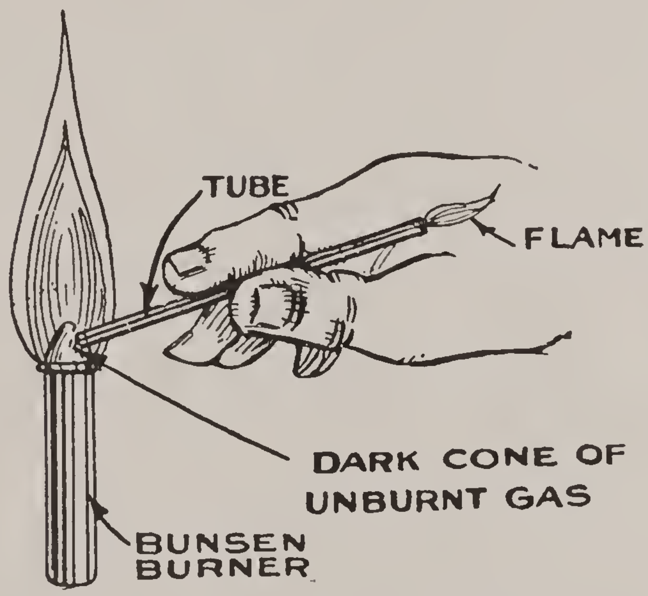

FIG. 140.—Proving the Dark Cone to be Unburnt Gas.

How to Make Colored Flames. To produce beautiful colored flames, get a piece of *pumice stone*, which is very porous and non-combustible, and fashion it into a ball about 1 inch in diameter; now fasten an iron wire around it, then dip it into any one of the following solutions and hold in the flame of an alcohol lamp or, better, a Bunsen burner, and the salt of which the solution is formed will give a characteristic color.

The solutions are made by dissolving the salts in water (H_2O), and these should be quite strong. Strontium chloride ($SrCl_2,6H_2O$) will give a bright red flame; calcium chloride ($CaCl_2$) will give a reddish-orange flame; copper chloride ($CuCl_2$) will give a bluish-green flame, and sodium chloride ($NaCl$) will give a brilliant yellow flame.

To see the colored flames to the best advantage, you must burn the salts in a dark room. A most curious effect is produced by burning sodium chloride ($NaCl$) so that the light from it will shine in the faces of the spectators, giving them a ghastly appearance. To prevent the solutions from dripping into the tube of the burner and so stopping it up, it is a good plan to lay it on its side, as shown in Fig. 141.

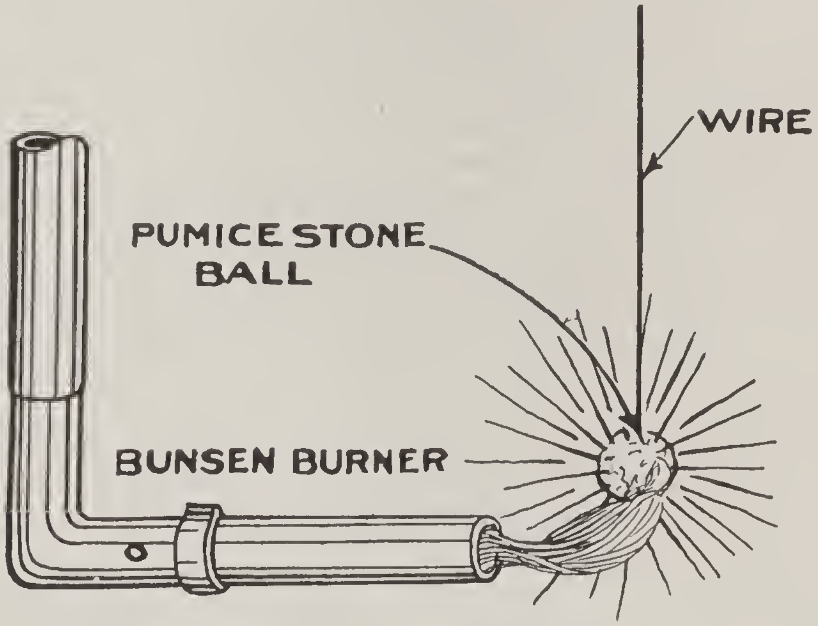

FIG. 141.—Production of Colored Flames.

How to Make Charcoal. Our sun is the original source of all the light, heat, and power we have here on earth, as well as everything else, and hence the energy stored up in all our fuels has come from it. For instance, the light and heat of the sun make plants grow, and these when large enough form trees, and are composed of wood.

In turn, wood is largely made of carbon (C), and by heating it in an enclosed space so that little or no air can get to it the gases are forced out of it and nearly pure carbon (*C*) is left behind. Now when carbon (*C*) is made to burn in oxygen (*O*), and the air supplies the latter, a very hot, flameless fire results.

To make a little charcoal (*C*), drive three sticks, each about 18 inches long into the ground, about 2 inches apart, then build up a conical pile of wood around them and leave a little space between the sticks, as shown in Fig. 142. This done, plaster it all over with mud so that the air cannot get through it to the sticks and then make a dozen holes, about 1 inch in diameter, through the mud covering, around the base, and also let the top remain open. This is so that when you light the wood the gases will slowly burn out of it and yet not burn the carbon (*C*) there is in it.

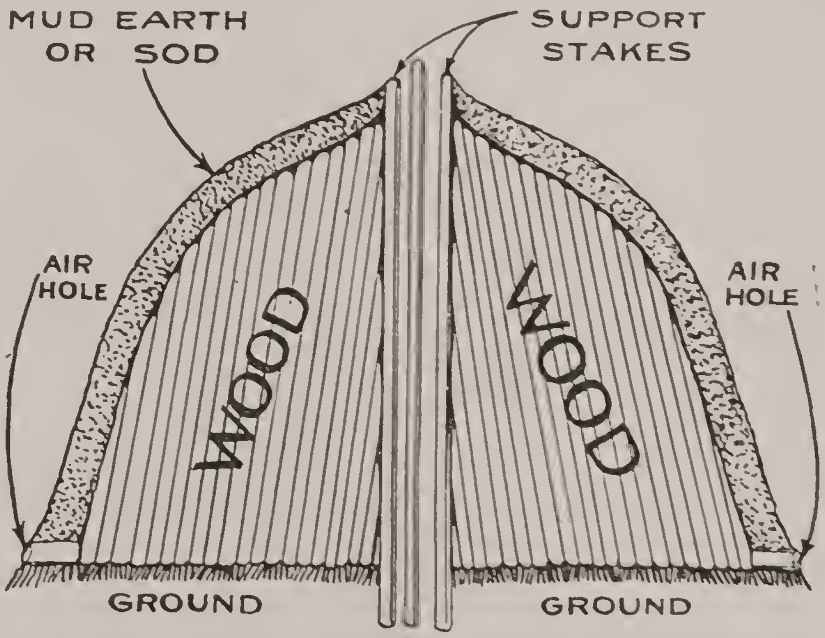

FIG. 142.—Cross-Section of a Charcoal Kiln.

How Charcoal is Made. When the wood begins to burn, the used gases pass out of the kiln and you will see them as a thick black smoke. After 24 hours or so, all the gases in the wood will be burnt out and only charcoal (*C*), which is an impure kind of carbon (*C*), is left behind. The substances in the wood that will not burn are left behind as ash.

What Coal is. When trees and other plant matter are covered over with sand or clay, as the great forests were in the pre-historic ages, so that the air cannot reach them, they decompose, that is, they *oxidize*, and the water (H_2O) and gases and oils that are in them are set free, and the matter that is left behind is called *coal* (*C*).

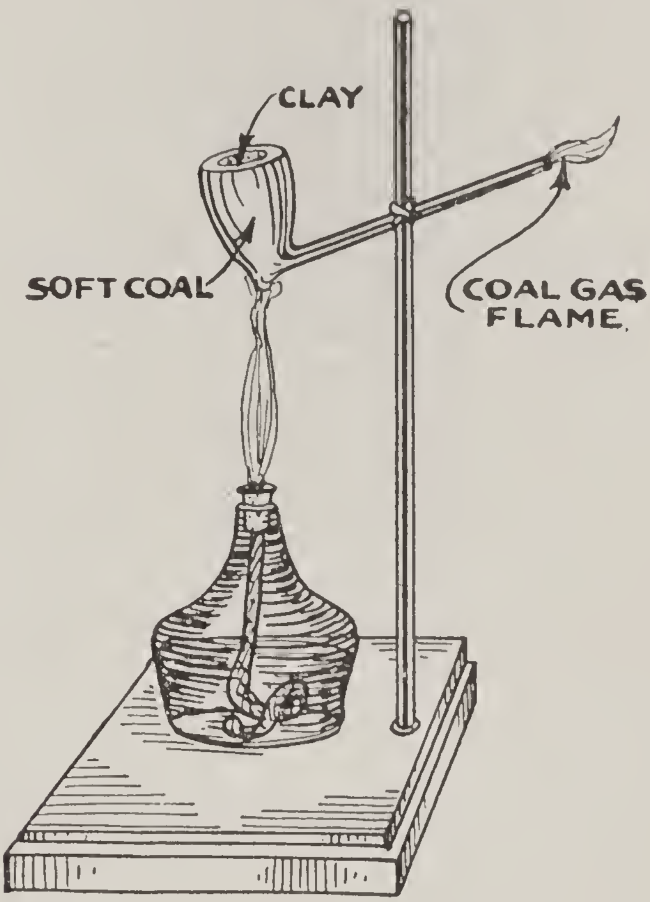

FIG. 143.—A Miniature Gas Works.

Now there are two kinds of coal (*C*), and these are *bituminous*, or *soft*, coal, and *anthracite*, or *hard*, coal. Bituminous coal is coal that still contains large amounts of hydrogen

(H) and oils of various kinds, and, hence, this kind is used for making illuminating gas. The coal is put into closed retorts, and after the gases and oils have been driven out of it by heat, there remain behind *coke* and *coal-tar*, and from the latter, dyes, perfumes, and medicines are made. Anthracite coal is nearly pure carbon (C), and it burns without flame, makes very little smoke, and leaves but a small amount of ash behind; it is, therefore, the most suitable kind of coal for heating purposes.

How to Make Coal Gas. Take a clay pipe and fill it with powdered soft coal and then close up the mouth of it with a piece of clay. Now heat the pipe in the flame of your alcohol lamp, or, better, because it is hotter, your Bunsen burner, and the hydrogen (H) will be driven out, and this you can light at the end of the stem, as shown in Fig. 143. When all the gas has passed out, you will find a little lump of hard, black porous matter in the pipe bowl, and this is coke, while the sticky substance that remains is coal-tar.

CHAPTER XII.

HOW TO MAKE PHOTOGRAPHS

THE art of photography is made up of three parts: light, optics, and chemistry. This may be explained by saying that light either coming directly from an object or reflected by it and made to pass through a lens of the proper kind, will form an image of the object on a flat surface, and if this is chemically prepared, the image can be fixed there and a picture, or *photograph*, as it is called, is thus made.

What Light is. In order to know how light acts, you must know something about its nature. In Chapter XI I told you that when a substance burns it gives out light; that the vibrating molecules of a burning substance set the ether into motion, and that it is by, in, and through the latter that *light-waves* travel. Now the two following analogues will make clearer what light and light-waves are: First, you have often noticed that when you throw a stone into a pool of still water, little ring-like waves, or circular ripples, will be formed around the place where the stone has struck the water, and these will expand until either their energy is used up or they are stopped by the shore. In other words, the stone sends out *water-waves*, as shown in Fig. 144.

Now, to go a step farther, if you strike a bell it *vibrates*, that is, the rim of it moves rapidly to and fro, as shown in

Fig. 145. These rapid movements are imparted to the air and set up waves in it, and while these are really air-waves, they are called *sound-waves*. These waves spread out in every direction and keep on expanding until they either strike some object and are reflected by it, or their energy is used up by in overcoming friction and other resistances.

Finally, if you ignite a substance that will burn, as, for instance, a candle, the heated molecules given off by it will be thrown into exceedingly rapid *vibration*, that is, a rapid to-and-fro motion, and these will set up waves in the ether which are called *light-waves*, see Fig. 146, but which are really only *ether-waves*. Like water-waves and sound-waves, light-waves are radiated in every direction, but normally travel in straight lines, and they keep on going until they are stamped out by the resistance they meet.

Fig. 144.—How a Stone Sends Out Water-Waves.

How Light Acts. Suppose you closed your eyes and that you held a string with a cork tied to it, and resting anywhere on the surface of the pool of water into which a stone was thrown. When the water-waves were sent out by the impact of the stone you would be able to sense their presence by the pull of the string every time the wave made the cork bob up and down, and the sensation of touch would by carried by the afferent nerves to your brain.

So in a like, but very much more refined way, wherever you may be, as long as you are within earshot of the sound-waves sent out by a bell, they will impinge on your ear, and the auditory nerves will convey them to your brain, where the sensation of sound is set up. Likewise, wherever you may be within range of the light-waves sent out, or reflected, by an object, the lens of your eye will form an image of it on the retina, and the optic nerves will transmit it to your brain, where the sensation of light, of form, and of color is produced.

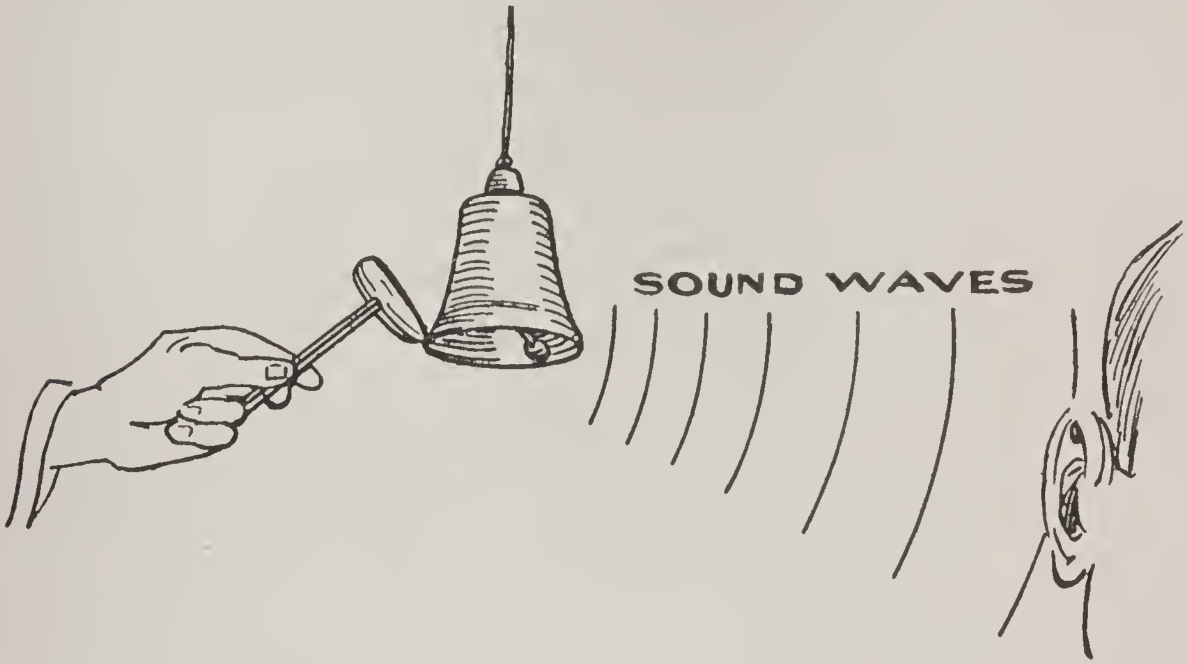

FIG. 145.—How a Bell Sends out Sound-Waves.

Now light not only acts on the eye so that we are able to see the images it forms, but it has a decided action on the growth of plants in that it builds up compounds in and for them; thus it makes the green coloring-matter of plants, called chlorophyl, and the action of light on this compound forms formaldehyde (CH_2O), which is a gas with a stifling odor, and this in turn is converted into sugar($C_{12}H_{22}O_{11}$). Not only plants but animals must have light in order to grow and, hence, its action on these bodies is to build up their tissues. While the action of light on plants and animals cannot be seen, there are compounds that break down, that is, they are decomposed, when exposed to it, and the effects of these are very readily observed.

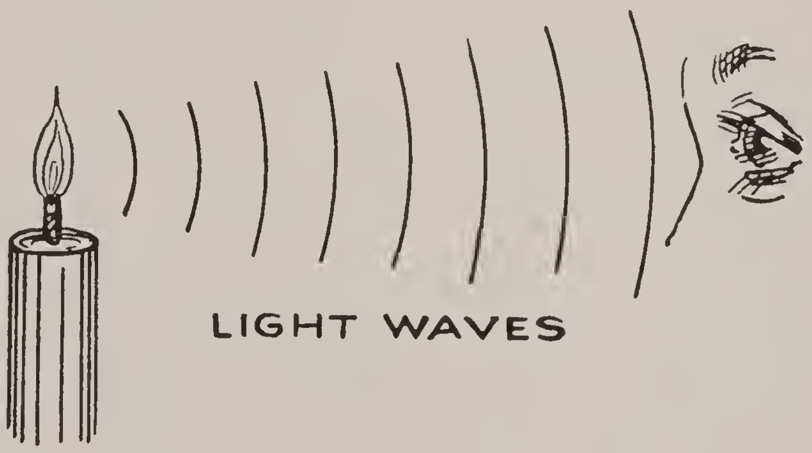

FIG. 146.—How a Candle Sends out Light-Waves.

How Light Acts on Silver. Of all the compounds that light has been found to act upon, those formed of silver (Ag) are the most sensitive, and for this reason they are used in photography. Of these salts, silver nitrate ($AgNO_3$), silver chloride ($AgCl$), and silver bromide ($AgBr$) are the most easily affected. Silver nitrate ($AgNO_3$) was the first salt that was found to break down under the action of light, and then followed silver chloride ($AgCl$), which is still more sensitive, and, finally, silver bromide ($AgBr$), which is the most sensitive of all.

How to Make Silver Nitrate. Silver nitrate ($AgNO_3$), which used to be called *lunar caustic*, is the starting-point for making the other salts of silver (Ag). To make a few crystals, put 2 fluid ounces of pure water (H_2O) in a beaker, then add ½ fluid ounce of nitric acid (HNO_3) to it and drop a bit of pure silver (Ag) the size of a dime into the solution. Stir it with a glass rod, and when the silver (Ag) has dissolved let it stand and crystals of silver nitrate ($AgNO_3$) will be formed, nitric oxide (NO) gas will pass off, and the liquid left behind will be water (H_2O) thus:

$$\underset{\text{Silver}}{Ag} + \underset{\text{Nitric acid}}{HNO_3} + \underset{\text{Water}}{H_2O} = \underset{\substack{\text{Silver}\\\text{nitrate}}}{AgNO_3} + \underset{\substack{\text{Nitric}\\\text{oxide}}}{NO \uparrow} + \underset{\text{Water}}{H_2O}$$

Experiments with a Silver Nitrate Solution. Nearly fill a clean 2-ounce bottle with distilled water (H_2O), drop in the crystals of silver nitrate ($AgNO_3$), put in the cork, and shake until the crystals are dissolved. If, now, you will place the bottle where the light of the sun will fall on it, no chemical action will take place and the solution will remain colorless. Now take a sheet of paper, pour the solution over it and expose it to the light of the sun, and you will find that the light will quickly act on it and turn it a brown color.

The question is why will not the light act on the solution when it is in the bottle as it does when it is spread out on the paper. The answer is because light will decompose the salts of silver (Ag) only when the latter is in contact with *organic matter*. By organic matter is meant plant or animal matter that is living or has once lived. Paper, as you know, is formed of cellulose ($C_6H_{10}O_5$), and it is of this compound that plants are largely built up.

How to Make Silver Chloride. Dissolve 1 teaspoonful of sodium chloride ($NaCl$), that is, common table salt, in a test tube full of water (H_2O), then dissolve the same amount of silver nitrate ($AgNO_3$) in a test tube one-fourth full of water. After the salts have dissolved, pour the solutions of both test tubes into a beaker, as shown in Fig. 147, and stir them together with a glass rod, and a milky-white precipitate will be thrown down, which is silver chloride ($AgCl$). The reaction is called a double decomposition, in which the silver (Ag) of the silver nitrate ($AgNO_3$)

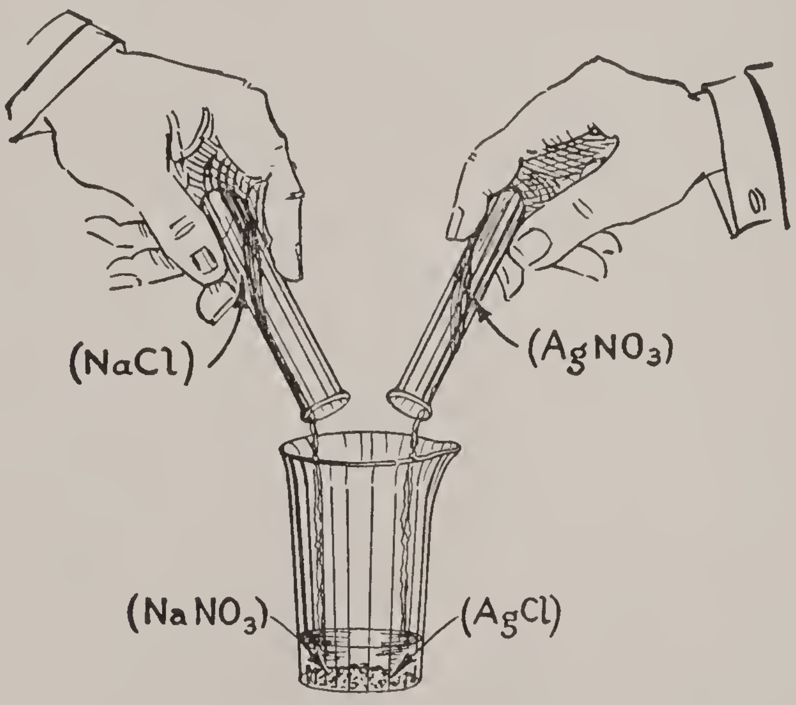

FIG. 147.—How Silver Chloride is Made.

changes places with the sodium (Na) of the sodium chloride ($NaCl$), and it may be expressed thus:

$$\underset{\substack{\text{Sodium}\\\text{chloride}}}{NaCl} + \underset{\substack{\text{Silver}\\\text{nitrate}}}{AgNO_3} = \underset{\substack{\text{Sodium}\\\text{nitrate}}}{NaNO_3} + \underset{\substack{\text{Silver}\\\text{chloride}}}{AgCl} \downarrow$$

The next thing to do is to put a sheet of filter paper in a glass funnel and then set this in a bottle; now pour the solution with the precipitate into the funnel. The solution will run through it and the precipitate will remain behind.

Action of Light on Silver Chloride. Put a tablespoonful of water (H_2O) in a test tube, and after scraping the silver chloride ($AgCl$) from the filter paper spread it over a sheet of unglazed paper and let it dry. If, now, you will expose it to the light of the sun you will see that it turns a purplish color first, then gets brown, and finally black.

Evidently the light has produced a change in the silver chloride ($AgCl$), and, in truth, it has acted on it in such a way that the compound has broken down into the two elements of which it was formed, namely, silver (Ag) and chlorine (Cl), as the following equation shows:

$$\underset{\text{Silver chloride}}{AgCl} + \underset{\text{Light}}{Light} = \underset{\text{Silver}}{Ag} + \underset{\text{Chlorine}}{Cl} \uparrow$$

The chlorine (Cl), which is a gas, passes off, and the very fine brown, or black, film that remains behind is formed of minute particles of pure silver (Ag). It is this action of light on silver compounds that makes it possible to take a picture on a glass plate, a celluloid film, or a paper sheet; there are, however, other operations necessary, the chief one being to fix the picture so that it will not fade out, and this will be described presently.

How to Make a Pinhole Camera. To understand how light forms a picture, or *image*, as it is more properly called, of an object, we shall have to leave the *chemistry* of it for the moment and get into the *physics* of it. The simplest way is to make a *pinhole camera*, by means of a pair of open-end, rectangular pasteboard cases, each of which is, say, 4 inches wide and high and 6 inches long, so made that one will slide snugly into the other, as shown in Fig. 148.

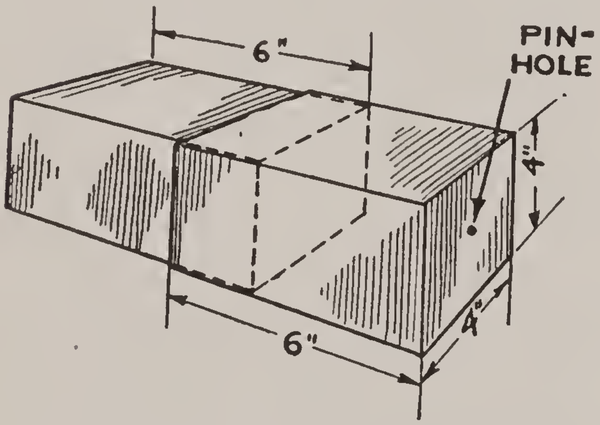

Fig. 148.—How to Make a Pinhole Camera.

Now glue a thin disk of cardboard over one end of the larger case and make a pinhole in the center of it, and then secure a sheet of oiled tissue paper over one end of the smaller case and slide them together. Your pinhole camera is then ready to use. Hold it in a line with the object the image or picture of which you want to see on the *screen*, as the oiled tissue paper is now called, and *focus* it, that is, slide the smaller case in or out until the image on the screen is as sharp as you can get it.

How the Camera Works. You will observe that, curiously, the image on the tissue paper is *reversed*, that is, it is upside down, but the reason for this will be clearly understood by a look at the diagram shown in Fig. 149. Now light-waves travel in straight lines and they are sent out in every direction from every point of a candle or other object, but of all the waves sent out from a particular point, as for instance the one marked A, only those will go through the pinhole, B, that are in a straight line with it, and then they pass on to the screen, where they strike it at C. In the same way, only the waves from the point marked D can get through the pinhole, B, that are in a straight line with it, and these impinge on the screen at E; and this is true of all other parts of the candle or other object.

How a Real Camera is Made. A real camera differs from the one just described in that it has a lens instead of a pinhole in the front part of it, and either a *plate-holder* to hold the sensitized glass *dry plate*, or a *roll-holder* to carry the spools, where *films* are used. A cross-section of a real camera is shown in Fig. 150. By using a lens in a camera, a great deal more light can be got through it than through a pinhole, and this makes for speed of exposure, and, further, and what is equally important, the image is very much more clearly defined.

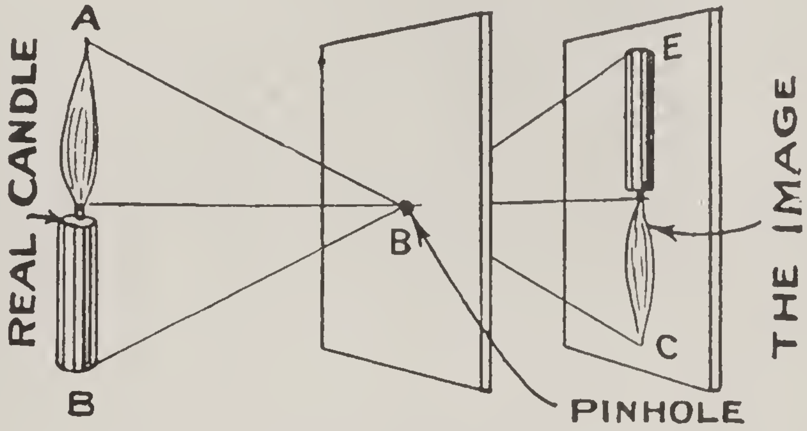

FIG. 149.—How the Image is Reversed.

How Dry Plates and Films are Made. To make a dry plate or a film so that it will be sensitive to the light and free from pinholes and spots is an expert's job, and he must have a specially equipped laboratory for the purpose. However, I will tell you how it is done and you can try to make them just as I did when I was a boy of your age, only I had the decided advantage of having worked for a dry-plate manufacturer.

Now just as silver chloride ($AgCl$) is more sensitive to light than silver nitrate ($AgNO_3$), so silver bromide ($AgBr$) is more sensitive than silver chloride ($AgCl$). You can make enough silver bromide ($AgBr$) *emulsion* to coat a dozen 4 by 5 glass plates in this way: put 1 ounce each of silver nitrate ($AgNO_3$) and ammonium bromide (NH_4Br) in a beaker, with enough water (H_2O) to dissolve them; now put 2 ounces of clear gelatine, which is an organic compound, in the beaker with enough water (H_2O) to cover it; gently heat them over the flame of your alcohol lamp until

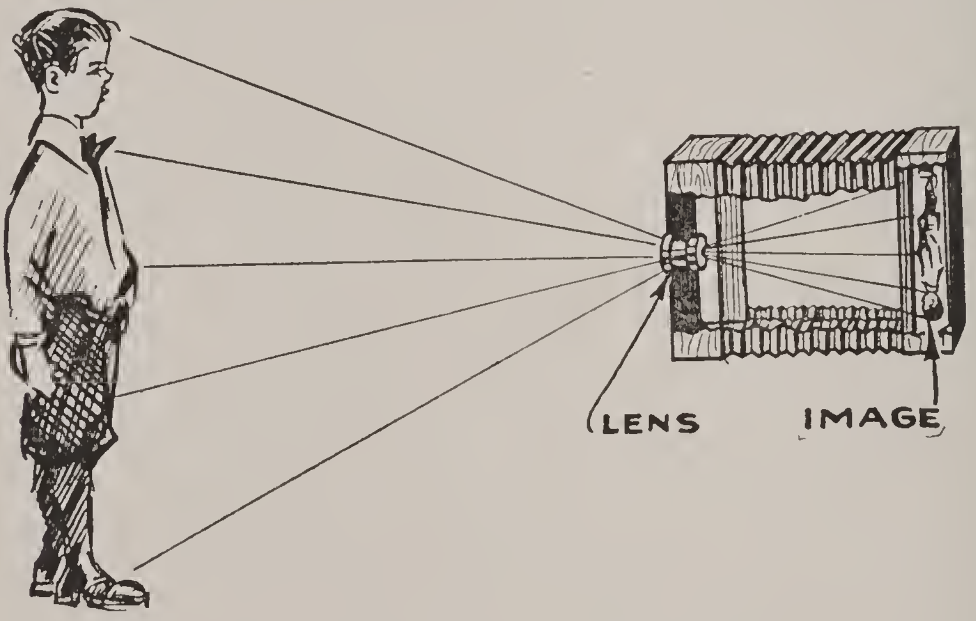

Fig. 150.—How a Real Camera Works.

the silver nitrate ($AgNO_3$) and ammonium bromide (NH_4Br) react on each other and form ammonium nitrate (NH_4NO_3) and silver bromide ($AgBr$), which is a double decomposition, thus:

$AgNO_3$	+	NH_4Br	=	NH_4NO_3	+	$AgBr$
Silver nitrate		Ammonium bromide		Ammonium nitrate		Silver bromide

You will easily know when this reaction has taken place, for the silver bromide ($AgBr$) will form in little drops, or granules, all through the gelatine solution, or emulsion,[1] as it is now called. Now let the latter get cool, becoming about as thick as jelly, and cut it up into bits about ⅛ inch square. The next thing to do is to soak the emulsion in water (H_2O) over night in order to wash all the ammonium nitrate (NH_4NO_3) out of it, leaving only the silver bromide ($AgBr$).

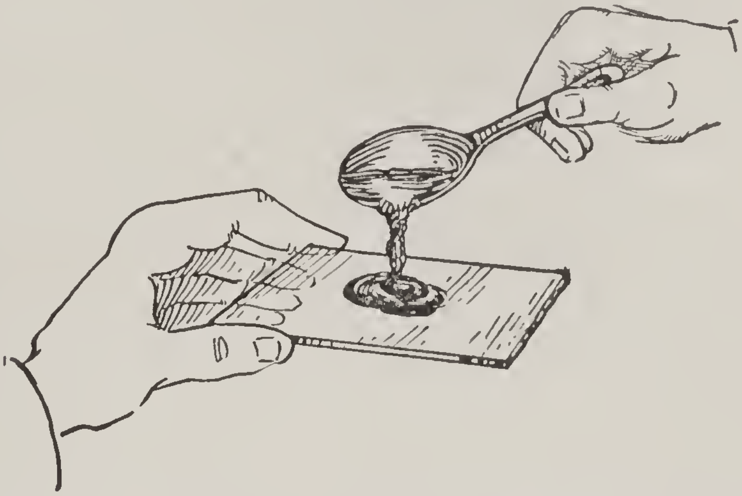

FIG. 151.—Coating the Plate with Silver Emulsion.

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[1] Strictly speaking, an emulsion is a liquid in which the solid particles of some other substance are held in suspension, that is, evenly distributed through it.

This done, melt the emulsion and then hold a perfectly clean glass plate in one hand, as shown in Fig. 151, pour a tablespoonful of it in the center, quickly tilt the plate so that the emulsion will flow all over it, then lay it on a perfectly level surface and let it dry over-night. All these operations must be performed in a dark room illuminated only by a very feeble red light, and while there should be a good current of air circulating in the room, it must be absolutely free from dust. Films are made by coating long strips of celluloid with the same kind of silver bromide ($AgBr$) emulsion as that described above, but it takes machinery of a special kind to do this evenly.

How a Picture is Made on a Dry Plate or a Film. When you make a picture with your camera, the image formed by the lens falls on the dry plate, or film, and the light instantly acts on the silver bromide ($AgBr$) on the gelatine surface in proportion to its intensity, and decomposes the silver bromide ($AgBr$) into particles of pure silver (Ag) and bromine (Br). The result is that the silver (Ag) remains on the plate or film as a brown powder and the latter is set free, thus:

$$\underset{\text{Silver bromide}}{AgBr} + \underset{\text{Light}}{Light} = \underset{\text{Silver}}{Ag} + \underset{\text{Bromine}}{Br}$$

This reaction, which may take place in the 1/1000 part of a second, or less, cannot be seen in the sensitized surface, and in order to bring out the picture, the plate must be *developed*, to dissolve and wash away those parts of the silver bromide ($AgBr$) which the light has not acted on.

How to Develop a Dry Plate or a Film. To develop a plate, or a film, you must soak it in a solution called a *de-*

veloper, and this you can make by dissolving 30 grains of hydroquinone ($(KO)_2CrH_4$), 10 grains of metol, 350 grains of sodium sulphite (Na_2SO_3), 350 grains sodium carbonate (Na_2CO_3), and 5 grains of potassium bromide (KBr), and 10 ounces of distilled water (H_2O). In making up this developer, use only the very best chemicals, and see to it that the sodium sulphite (Na_2SO_3) and sodium carbonate (Na_2CO_3) are good clear crystals.

Now when you soak the exposed plate, or film, in this developer, the gelatine is softened by it and the bromine (Br) that has been separated from the silver bromide ($AgBr$) by the action of light is absorbed by the developer, and this leaves the pure silver (Ag) behind. As the development goes on, you can see the picture slowly "come up," that is, come into view — a most fascinating process — as the contrast grows greater between the parts which the light has affected and those which it has not affected. The parts, however, that were *white* of the object which was photographed will show in the developed plate as *black*, because the silver (Ag) that has remained behind is black, and, oppositely, the parts of the object that were *black* will show as *white*, for here the silver bromide ($AgBr$) was not affected. In other words, the black and white parts on the plate, or film, are just the reverse of those of the object that was photographed, hence, the plate is now called a *negative*.

How to Fix the Picture. If the developed plate, or film, should again be exposed to the light it would decompose the remaining silver bromide ($AgBr$), and all of it would be decomposed into silver (Ag) and bromine (Br). To keep this action from taking place when the picture has reached the proper stage of development, it must be *fixed*, as it is called, and, naturally, thus must also be done in a dark room.

To fix the plate, or film, you must soak it in a *fixing bath*, which is simply a solution that you make by dissolving ½ pound of sodium thiosulphate, ($Na_2S_2O_3$), often incorrectly called *hyposulphite*, or *hypo*, for short, in ½ pint of boiling water (H_2O), and then adding another ½ pint of cold water (H_2O) to it. Keep this in a corked bottle until you want to use it. When you are ready to *fix* the picture, put the plate, or film, in a glass tray three-fourths full of the hypo fixing bath and let it stay there until all the silver bromide ($AgBr$), not acted on by the light, is dissolved out; you will know when this action has taken place, for the opaque whiteness of the plate disappears, and you will be able to see the transparent picture by holding the plate, or film, to the light.

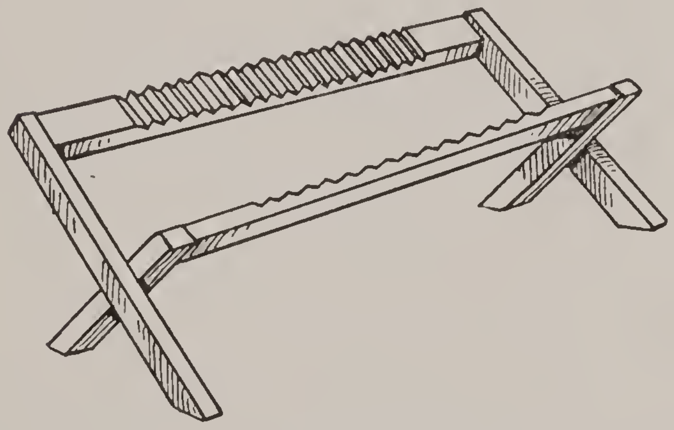

FIG. 152.—A Negative-Rack.

Let the negative remain in the fixing bath for ½ hour or more so that every particle of the silver bromide ($AgBr$) may be dissolved, and then wash it for an hour under running water (H_2O), or in many changes of it, in order to remove all the hypo with which the gelatine coating is saturated and which if not removed will stain the negative. After washing the negative, set it in a rack, see Fig. 152, and let it dry slowly in a cool place where there is a good circulation of air.

How to Make a Print from a Negative. Now while the white and black parts of the picture on the negative are just the reverse of what they were of the object you photographed, you can make a *positive*, or as many as you want, on paper or on glass. When positive copies are made on paper they are called *prints*, and when they are made on glass they are called *transparencencies*, if they are to be viewed by the light shining through them, or *lantern slides*, if they are to be projected on a screen.

Kinds of Printing Papers — Silver Papers. There are different kinds of papers used for making photographic prints, but all are coated with either a nitrate, a chloride, or a bromide silver compound. Those coated with silver nitrate ($AgNO_3$) are slow printing papers and must be exposed to the sunlight, hence, they are called *printing-out* papers. There are two kinds of printing-out papers, the first of which is known as *silver paper*. This is coated with albumen, which is *white-of-egg*, and therefore an organic substance, and then with a solution of silver nitrate ($AgNO_3$); the second is called *solid paper*, and this is coated with gelatine, also an organic substance, and then with a silver chloride ($AgCl$).

How to Make a Print. To make a print, you need a *printing-frame*, as shown in Fig. 153. Take the back out of it, lay the negative in the frame with the film side up, that is, toward the back, and lay the sheet of sensitized paper on it, with its film side down, that is, next to the negative. Now put the back in the frame and clip the ends of the springs, which are pivoted to the back, under the catches that are fixed to the frame.

If you are making a silver or solid print, set the frame out of doors so that the sunlight will fall directly on the negative. From time to time take the frame into a more subdued light and unclip one of the springs. Then you can lift up half of the back (it is hinged together) and look at the print to see how it is coming on. The print when ready to be taken out of the frame will be a positive, for the light that goes through the clear parts of the negative will turn the paper brown or black, and, conversely, where the negative is black, the light cannot get through, and so the paper remains white.

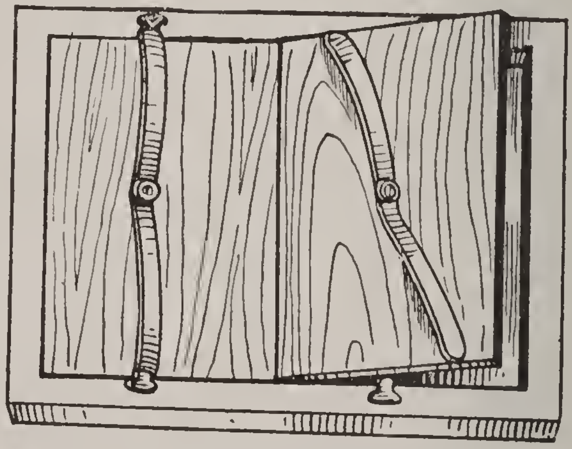

FIG. 153.—A Printing-Frame.

How to Tone the Print. The appearance of a silver print is never very pleasing as it comes from the frame, and to give it a soft rich color you must *tone* it. This you do by putting it in a tray that contains a *toning solution*, as it is called; it consists of two solutions which you make up as follows: put 7 drams of distilled water (H_2O) in a small bottle and dissolve 7 grains of *auric chloride* ($AuCl_3$), or *gold chloride*, as it is commonly called, in it and label it, *Solution No. 1.* Then put 5 ounces of water (H_2O) in another bottle and dissolve 220 grains of ammonium sulphocyanide ($(NK_4)_2HSCN$) in it and label it, *Solution No. 2.* When you are ready to tone the print, put 2½ ounces of water (H_2O) in a tray and then add 1 dram of Solution No. 1; stir it with a glass rod slowly, and then put in 1 dram of the Solution No. 2; let it stand for 10 minutes, and it will then be ready for use. Now put the print in this solution and keep turning it over constantly until a rich deep-brown color is reached; next, wash it in two or three changes of water (H_2O), and then fix it for 15 minutes in a fixing bath made of ¾ ounce of sodium thiosulphate ($Na_2S_2O_3$) dissolved in 5 ounces of water (H_2O), and in which you have stirred a drop or two of liquid ammonia (NH_4OH).

After fixing the print for ½ hour, wash it in running water (H_2O) for an hour or so, or in numerous changes of water (H_2O). Finally, dry the print, and you will have a finished photograph.

How to Make a Velox Print. About 20 years ago a paper was introduced under the trade name of *velox*. This new kind of paper, which is coated with a silver bromide ($AgBr$) emulsion like plates and films, gives a beautiful black-and-white print. The great advantage of using it lies in the fact that it can be printed by gas-light — hence in England it is called *gas-light paper* — and while it must be developed, like a plate or a film, this does not take anywhere nearly the length of time that printing and toning a silver print does. Because of these advantages it soon found favor with both amateur and professional photographers, and it was not many years before the silver print was entirely supplanted by it except for commercial art work.

How to Make and Use Blue Paper. This paper is by all odds the cheapest and simplest kind to make and use, since a salt of iron is employed for sensitizing the surface of it, and it only needs to be washed thoroughly to bring out the picture and to fix it. For these reasons it is largely used by engineers and architects for making prints of drawings and plans, but you will find it gives you pretty prints of many objects and especially of marine views. While the paper is easy to make, still, owing to the poisonous nature of the chemicals employed, I would advise you to buy it ready-made.

Blue paper is made by dissolving ¼ ounce of green iron ammonio-citrate in 1 ounce of water (H_2O), and the same amount of potassium ferrocyanide ($K_3Fe(CN)_6$) in a like amount of water (H_2O). The two solutions are now mixed together and the surface of some unruled sheets of writing paper, or other well-sized paper, is coated over with it by means of a brush. It is then dried in a dark room, after which it is ready to be printed in the sunlight like silver paper. The only other operation is to wash it for ½ hour in running water (H_2O), or in many changes of it.

CHAPTER XIII.

THE WHITE MAGIC OF CHEMISTRY

THERE are a great many experiments in chemistry that have been used by magicians the world over during the last half-century, and these are as pleasing to-day as when they were just invented. In recent years, however, the knowledge of chemistry and of chemical processes has advanced to such an extent that the average spectator is not so easily deceived as he once was, but, curiously enough, even though he has an idea of how the tricks are done, in the last analysis, the effects are still quite as wonderful, for chemistry *is* magic. In this chapter I shall tell you how to perform enough startling experiments for a show that will last for half an hour or more.

Pouring Wine and Water from the Same Pitcher.— The Effect. Like a miracle of old, you pour from the same pitcher wine or water (H_2O) as the audience calls for it. On the table you have a clear glass pitcher full of water (H_2O) and half a dozen empty tumblers standing in front of it, as shown in Fig. 154. After a few remarks on making your own wine you ask the audience which it prefers, wine or water (H_2O), and you proceed to fill one of the tumblers with whichever beverage is called for.

When you have filled half of the tumblers with wine and half with water (H_2O), you pour them back into the pitcher and all will instantly change into wine, which you prove by filling up the tumblers. This done, you pour the wine back into the pitcher, and it is changed instantly into water (H_2O), as at the beginning, and you demonstrate the fact by filling up the glasses with it.

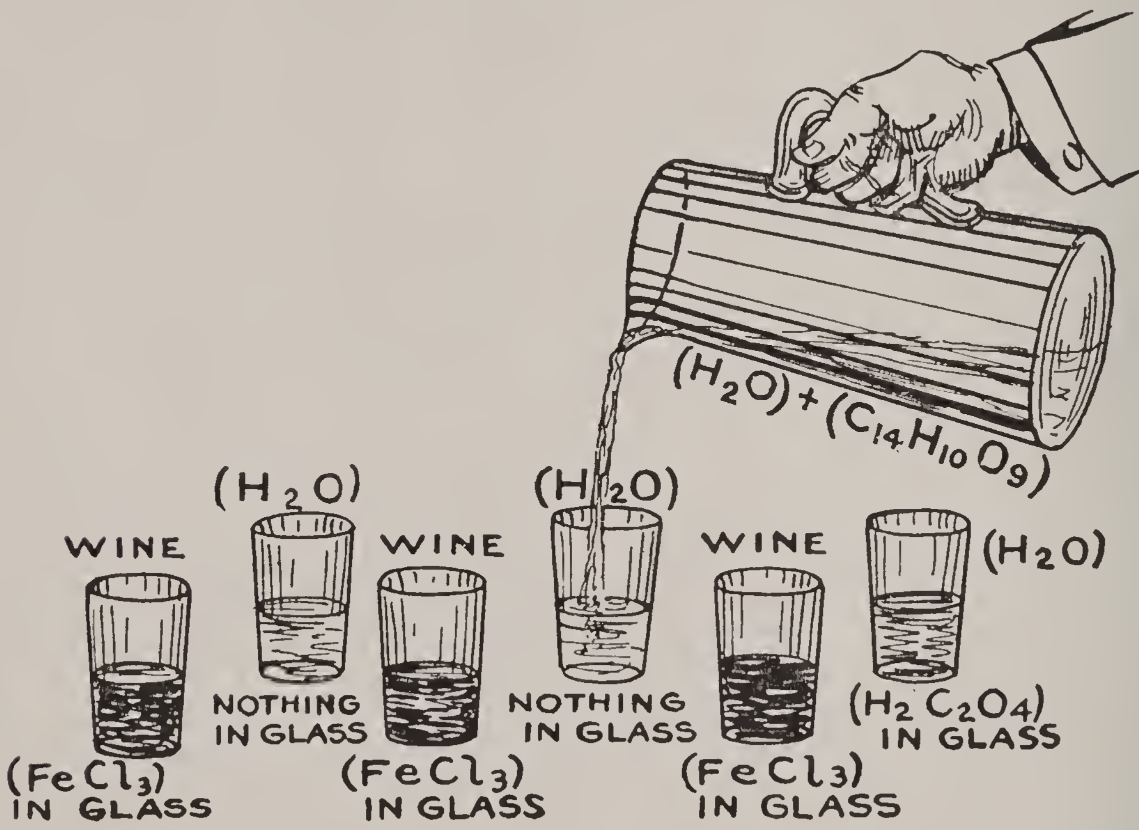

Fig. 154.—How Wine and Water are Poured from the Same Pitcher.

The Cause. It takes not the slightest skill to perform the trick. All you have to do is to dissolve 1 tablespoonful of tannic acid ($C_{14}H_{10}O_9$), which is a brownish powder made of nut-galls, in a pitcher of clean warm water (H_2O); now put ¼ teaspoonful of oxalic acid ($H_2C_2O_4$), which comes in needle-shaped white crystals, into one of the tumblers and pour on just enough hot water (H_2O) to dissolve them; finally, put 3 or 4 drops of *tincture of iron*,[1] which is ferric

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[1] You can get all these chemicals at a drug store.

perchloride ($FeCl_3$) dissolved in alcohol (CH_4O), into each of the other three tumblers. The two remaining tumblers have nothing in them.

You are now ready to do the trick and to ask the audience to say whether they want wine or water (H_2O). If wine is called for, you fill up one of the tumblers that has tincture of iron in it; if water (H_2O) is named, fill up one of the tumblers that has nothing in it; but in any event always fill the tumbler that has the oxalic acid ($H_2C_2O_4$) in it.

The instant the water (H_2O) from the pitcher comes in contact with the tincture of iron it will turn the color of wine. When you pour back into the pitcher the contents of the three tumblers it will color the water (H_2O) you have poured in from the other two tumblers, and you can then pour out all wine.

To change this back into water (H_2O), pour the oxalic acid ($H_2C_2O_4$) solution in the last tumbler into the pitcher first and then pour in the wine; the instant the iron of the latter comes in contact with the acid, a reaction takes place which precipitates the iron, and so leaves the water (H_2O) as clear at the end as it was at the beginning. The arrangement is clearly shown in Fig. 154.

Changing Water into Ink, and Vice Versa.—The Effect. You show a decanter half full of water (H_2O), as shown in Fig. 155, and one half full of ink, then cover each of them with a borrowed handkerchief and give them to two spectators who are some little distance apart, to hold. Now with a few magic passes you command the water (H_2O) to change into ink and the ink to change into water (H_2O), and when you pull the handkerchiefs from the decanters the audience will see that these transformations have truly taken place.

The Cause. This is a modification of the Wine and Water Trick described above, but instead of using the chemicals in a loose state you use them in tablet form, and these you can buy of dealers in magical apparatus and supplies. To change the ink in the decanter into water (H_2O), you need only to drop in an *acid tablet*, and to change the water (H_2O)

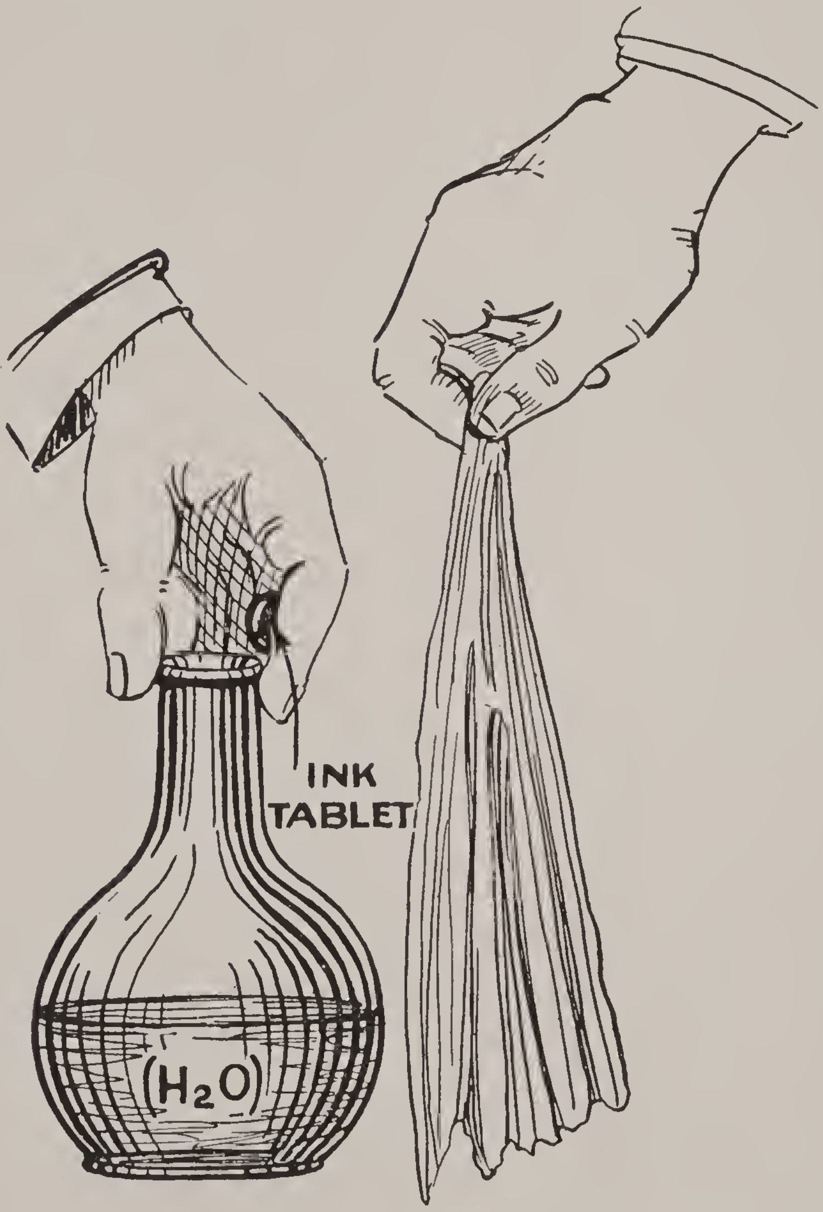

Fig. 155.—Changing Water into Ink.

to ink in the other decanter you simply drop an *iron tablet*, or *ink tablet*, as it is called, into it. To prevent the tablet from being seen by the audience, you clip it between your index and middle fingers, as shown in Fig. 156, and as you hold the decanter by its neck you drop the tablet in just as you throw the handkerchief over it with your other hand.

The Blushing Bride.— The Effect. For this trick you draw a picture of a beautiful girl — or if you can't draw a beautiful girl, get an artist friend to do one for you — and this you show to the audience. Then lay it on the table and rub it gently with your finger-tip, and when you show it again, the girl will be seen to be blushing like a sweet graduate of 17, or thereabouts. In a moment or two she will recover and the blush will disappear.

FIG. 156.—How the Ink Tablet is Held.

The Cause. Before you show the picture, paint either the cheeks or the whole picture with a solution made by mixing 1 tablespoonful of methyl alcohol (CH_4O) or wood alcohol, as it is called, in a like amount of water (H_2O) and add just enough phenolphthalein ($C_{20}H_{14}O_4$) to color it. The color will not show on the picture until the latter is brought into contact with the fumes of ammonia (NH_3). To do this, you dampen a sheet of blotting paper with liquid ammonia (NH_3), and it is on this that you lay the picture under the pretext of rubbing it.

The Magical Atomizer.—The Effect. You show half a dozen white feathers to your audience and then stick them into a frame or holder, as in Fig. 157. This done, you go among the spectators and spray them with *eau de Cologne* from an ordinary atomizer, just to prove that you really have perfume in it. You can ask them now to call out the colors they want you to make the feathers. One will say red, another blue, a third green, and so on, and as each color is named, you spray a feather with the atomizer, as shown in Fig. 158, and it instantly turns the color that has been called for.

Fig. 157.—The Feathers in Their Support.

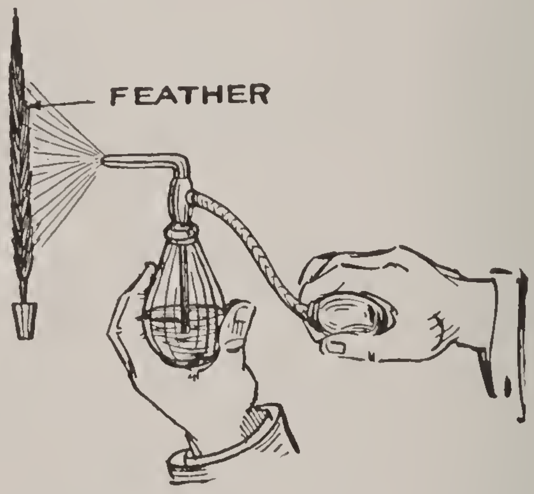

Fig. 158.—Spraying a Feather.

The Cause. Before beginning the trick fill the atomizer with methyl alcohol (CH_4O) and put just enough *eau de Cologne* in it to kill the odor of the latter and to make it smell like real perfume. The next thing to do is to dust each feather with a different-colored aniline dye, or *diamond dye*, which you can get at the drug store. Shake the feathers after you dust them with the dyes, and the particles left on them cannot be seen. Now the instant the alcohol (CH_4O) strikes the dye on the feather it dissolves it, and the feather becomes beautifully colored.

The Rainbow Liquid.—The Effect. This trick is called the *rainbow liquid*, for the very simple reason that from the way it acts it is clear you must have broken off the end of the rainbow in it. First you show a tumbler perfectly empty, of course, and then you fill it with water (H_2O) from another tumbler. As soon as you pour the water (H_2O) from the second tumbler into the first one, it turns green, not necessarily from envy, then it changes slowly to blue, this is transformed into violet, next it takes on a purple color, and finally it settles down to red. The transformation of one color into another without the glass being touched in any way is very mystifying.

The Cause. To do this trick, powder 1 tablespoonful of manganese dioxide (MnO_2) and 3 tablespoonfuls of potassium nitrate (KNO_3) in your mortar, and then put them into a sand crucible. Bring this mixture to a red heat in a stove without covering the crucible, and potassium oxide will be formed.[1]

When the compound is cold, put a few grains of it secretly into a tumbler, and then you are ready to do the trick. It is a good plan to experiment with different proportions of the potassium nitrate (KNO_3) and the manganese dioxide (MnO_2), and also to try warm water (H_2O) instead of cold.

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[1] This you can buy already prepared.

Breathing a Picture on Glass.—The Effect. You show a perfectly clean sheet of glass, say 4 by 5 inches on the sides, to your audience and let the members examine it as closely as they wish. Now breathe on it, and a well-defined picture will appear on the surface, as shown in Fig. 159. In a moment the picture will disappear even more mysteriously than it came.

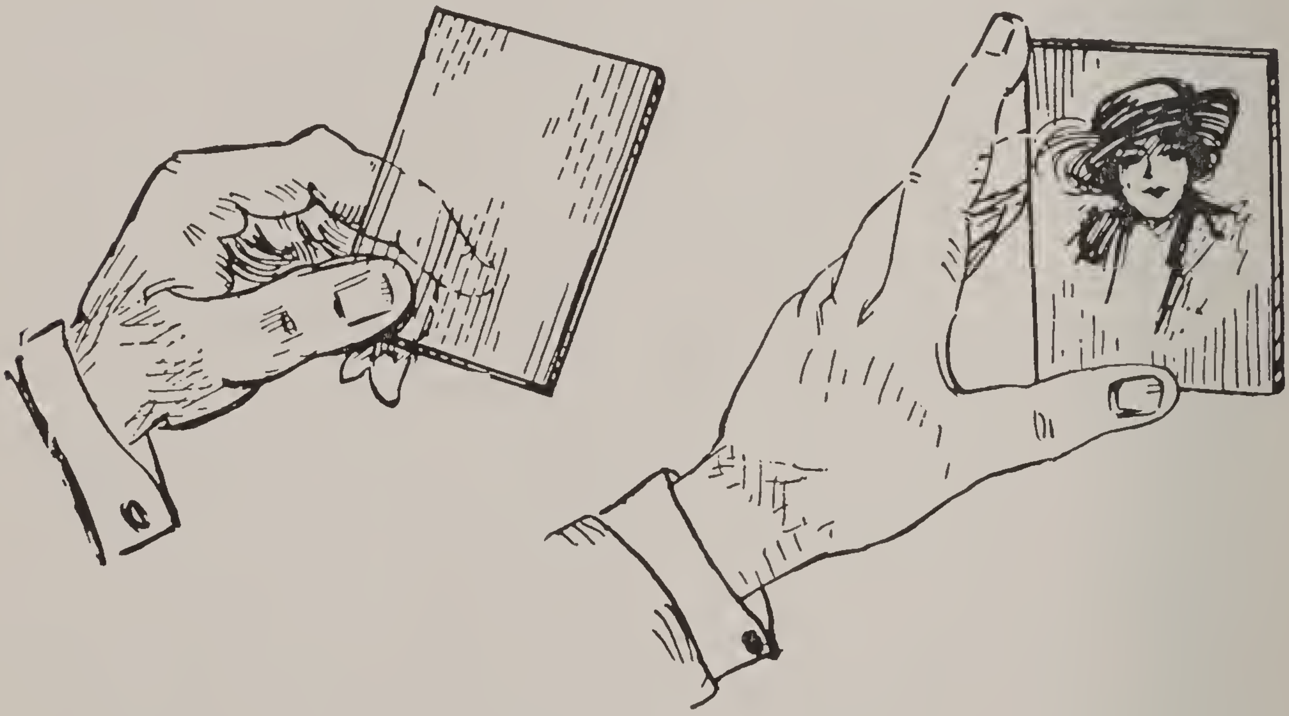

FIG. 159.—Breathing a Picture on Glass.

The Cause. This trick is simply an experiment with etched glass. To do it, get a sheet of glass that is perfectly clear and with a new steel pen draw a picture on it with hydrofluoric acid (HF), see Chapter VII. Let the acid remain on the glass for 8 or 10 minutes and then wash it off and dry the surface with clean cloth. The picture will be invisible even when the glass is examined closely, but it will be made visible the moment you breathe on it.

To get the right depth to the etching for it to be invisible when the glass is dry and yet stand out clearly when you breathe on it, you should make half a dozen of them and let the acid remain on each one a different length of time.

NOTE:— *Be very careful not to get any of the acid on your fingers.*

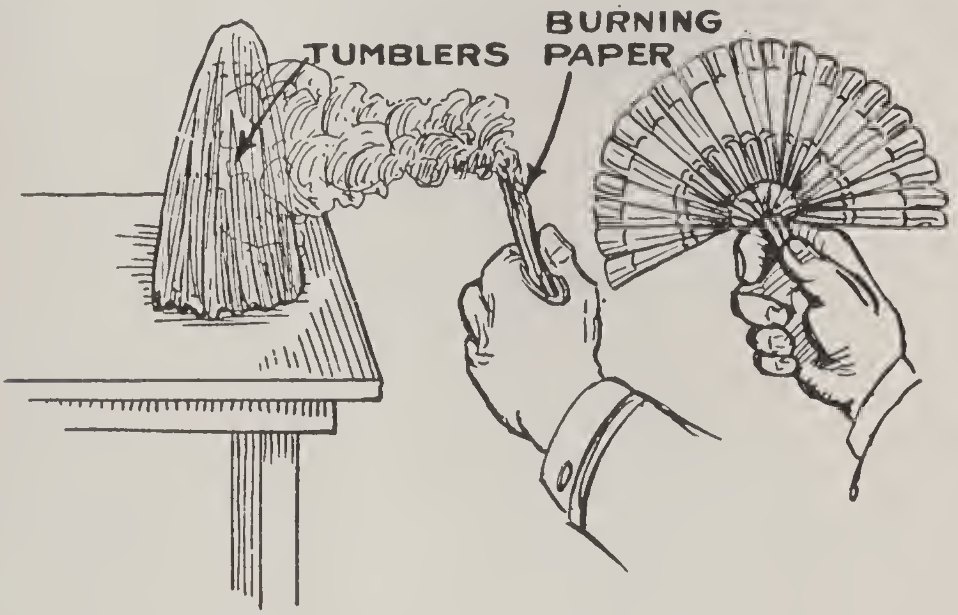

FIG. 160.—Passing Smoke Invisibly into the Glass Tumblers.

Passing Smoke Invisibly into a Tumbler.—The Effect. You show two empty glass tumblers to the audience, then place them mouth to mouth, throw a borrowed handkerchief over them and set them on a table or, better, let an assistant hold them. This done, you fold up a strip of paper and light it, and as the smoke rolls up and away from it you fan it toward the tumblers, as shown in Fig. 160, and explain as plausibly as possible how the atoms of smoke are wafted across the intervening space and on coming in contact with the tumblers pass through the pores of the glass and so find their way inside. To prove it, you remove the handkerchief, and the tumblers will be seen to be full of smoke, and on taking the top one off, the smoke will rise in a cloud, as shown in Fig. 161.

Fig. 161.—Showing the Smoke in the Tumblers.

The Cause. To do this simple but astounding trick, put a few drops of hydrochloric acid (HCl) into one of the tumblers and turn it rapidly round and round so that the acid will cover as much surface as possible. Now put a few drops of concentrated liquid ammonia (NH_3) in the other tumbler and turn it rapidly round.

In this way the acid and the ammonia will dry on the surfaces of their respective tumblers and you can show them as being perfectly empty. Now when you put the tumblers together, the fumes of the acid and the ammonia will come together, and form ammonium chloride (NH_4Cl), which has the appearance of real smoke.

Elixir Vitae, or the Artificial Production of Life.— The Effect. You show a few grains of coarse sand and drop them into a soup-plate filled with water (H_2O). Instantly they will become to all intents alive, and will move and whirl about like some water insects, as shown in Fig. 162. Now touch the surface of the water (H_2O) with the end of your wand, a lead pencil, or your finger, and they will lose their lifelike qualities and become as motionless as the bits of inert matter they were at first.

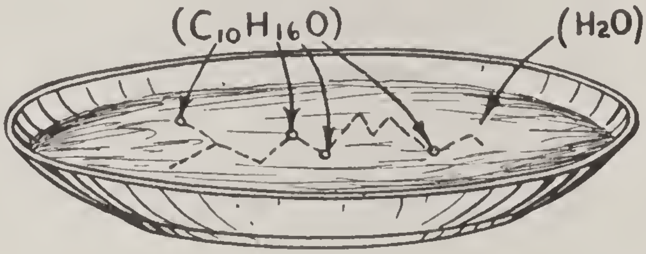

FIG. 162.—Elixir Vitae, or the Artificial Production of Life.

The Cause. The secret of this trick lies in the fact that the so-called grains of sand are really particles of camphor ($C_{10}H_{16}O$), and when this comes in contact with the water (H_2O) a reaction takes place in which hydrogen (H) is set free, and this makes the camphor ($C_{10}H_{16}O$) move about in a lively manner. The end of your wand, the pencil, or your finger, has a little grease on it, and when this comes in contact with the water (H_2O) it prevents it from acting on the camphor ($C_{10}H_{16}O$), and hence it gives up its false life.

How to Make Secret Writing Inks. Secret writing inks, or *sympathetic inks*, as they are generally called, are invisible when they are dry but become visible when they are acted on by light, heat, and various forms of electromagnetic disturbances.

A Heat Sympathetic Ink. Dissolve a very little cobaltous oxide (CoO) in hydrochloric acid (HCl), and deep red crystals of cobaltous chloride ($CoCl_2$) will be formed, or you can buy the latter compound already prepared, which is somewhat easier. Now dissolve these crystals in a little water (H_2O) and write with the solution just as you would with ordinary ink, but use a pink-colored paper. The cobaltous chloride ($CoCl_2$) ink will become invisible as soon as it dries, but to read what you have written it is only necessary, to warm the paper, and the ink will take on a blue color; as soon as it is cold it will take on a pink color.

How the Ink Works. The crystals of cobaltous chloride ($CoCl_2$) have a great attraction for water (H_2O). Now when the ink made of them dries on the paper, minute crystals of the compound are formed and these attract the water vapor in the air, which turns them a slightly pinkish tint, and they are practically colorless on the background of pink paper. But when they are heated, the *water of crystallization* is driven out of them and they then turn blue, and so stand out in relief on the pink paper.

A Light Sympathetic Ink. Dissolve a small crystal of silver nitrate ($AgNO_3$), or *lunar caustic*, as it used to be called, in a test tube half full of water (H_2O) and write with it on a sheet of white paper, using a sharp toothpick for a pen. When the ink is dry, the writing will be invisible, and remain so as long as the paper is kept folded up and away from the light. But as soon as it is opened and exposed to the light of the sun, the writing will become visible, taking on a brown color at first, and then turning to a jet black.

How the Ink Works. When the light-waves of the sun fall on the paper, they partly decompose the silver nitrate ($AgNO_3$) and set free the nitrogen (N) and the oxygen (O) and leave a brown powder behind, which is nearly pure silver (Ag). When enough light-waves fall on it entirely to decompose the silver nitrate ($AgNO_3$), then a black powder is left, which is pure silver (Ag).

A Fluorescent Secret Ink. Dissolve some quinine sulphate ($C_{20}H_{24}N_2O_2$), the kind you take for colds, in a little water (H_2O) and use this as an ink to write with. When it is dry it cannot be seen, but if you will hold it up close to the sparks of an induction coil, the writing will appear to be of a blue-violet color.

How the Ink Works. When the light-waves strike certain substances, they are absorbed by the latter, followed by the emission of light-waves of a different and greater length, and this phenomenon is called *fluorescence.* The short, invisible ultra-violet waves that are set up by the sparks of an induction coil are absorbed by the quinine sulphate ($C_{20}H_{24}N_2O_2$), which then sends out longer waves and these produce wave-lengths that make violet light which can be seen.

How to Make Spirit Pictures.—The Effect. You show a dozen pieces of perfectly blank paper, about 1 by 2 inches on the sides, and after they have been examined you ask a lady (an unmarried one of course) to select one of them, in order that you may show her her future husband. When she has selected one, you dip the blank paper into a saucer of water (H_2O) and while it is still wet you place it on her forehead. On removing it, there will be seen a photograph of a handsome young man with lots of money, a wonderful career before him, and all that.

The Cause. First of all, you make a dozen small prints from an ordinary photographic negative of a handsome young man, etc., etc., or better, make each print from a different negative. The prints must be made on what photographers call *silver paper*, such as was universally used 25 years ago but which is now employed chiefly by commercial artists for enlargements. The present-day *solio paper* will not do, and the silver paper must not be of the kind called *self-toning*, either.

After having made the prints, fix them *without toning* in a 10-per-cent solution of sodium thiosulphate ($Na_2S_2O_3$), and then wash them thoroughly. This done, immerse them in a 5-per-cent solution of mercuric chloride ($HgCl_2$), commonly called *corrosive sublimate*, and the picture will quickly fade out and the paper will appear to be perfectly blank. Finally, wash the prints again and let them dry, and you are ready to make the spirit photographs or, rather, make the spirits make the photographs for you.

Just before you are going to do the trick, make a 5-per-cent solution of sodium thiosulphate ($Na_2S_2O_3$), and this will look just like ordinary water (H_2O). Now when you dip the apparently blank paper into the solution, it only takes a moment for the reaction to make the picture reappear, and to prevent the lady from seeing this process you hold it on her forehead. A very pretty trick.

The Materialization of Mysteria.—The Effect. In the language of the spiritualist, the word *materialize* means to bring forth a spirit in bodily form so that it can be seen. Because spirits are made of stuff as intangible as dreams, they can be seen only when they are luminous and, hence, only in the dark, and so for this extraordinary test in psychophysical phenomena you must have a perfectly dark room.

When you are ready to materialize *Mysteria*, have your audience seated in one end of the room, then turn out the lights and your *dark seance* is on. First, the spectators will see an uncertain ghostly light, like a will-o-the-wisp, close to the floor and near the other end of the room. And then this strange light, certainly not of this earth but mayhap of heaven above, begins to expand and at the same time to take on a more definite shape until it can be clearly seen to be that of the form of a girl. When she has been fully evolved, her face, beautiful beyond words, materializes from out of the ambient astral light, and grows so brightly radiant that her very features can be recognized.

She is none other then *Mysteria*, the beautiful spirit-bride, who has come back to the earth-plane and her mission is to put to shame the scoffers who disclaim a life hereafter. See! she rises from the floor and floats in the air as lightly as a bubble. Returning, she grows smaller and smaller and beautifully less until she can just be seen as a vaporish patch of light, and then she *dematerializes* before the very eyes of the spectators.

The Cause. It almost saddens me to tell you how *Mys-*

teria is materialized but since this is a book of living as well as of dead secrets I will give you the explanation. First of all, you need several *props*, as they are called in the show business, two rooms that can be made perfectly dark, and an assistant. To make the former, get some soft iron wire of about No. 10 or 12 gauge and fashion it into the outline of a girl, as shown in Fig. 163.

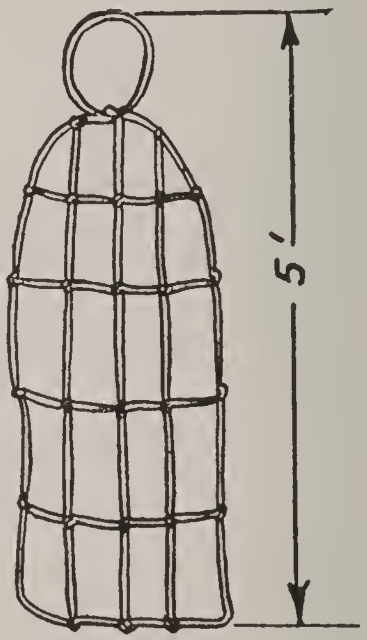

FIG. 163.—The Wire Frame.

Next, fasten a false face of a pretty girl to the top of it, and then paint this with *luminous paint* mixed with a little thin varnish. Luminous paint is made chiefly of phosphorus (P), so-called from two Greek words which mean *light* and *I bear*, and this element unites with the oxygen (O) of the air very slowly, and in so doing light is produced with practically no evolution of heat. You can buy from dealers in magical supplies luminous paint ready to use, with the proper varnish to thin it down.

Having prepared the face, take about 4 yards of cheesecloth, tack it to a wall or other flat surface and then paint it over lightly with the rest of the luminous paint, which you have thinned down with a quart of the varnish. When the paint is quite dry, make a simple one-piece gown of the cloth, like a night-dress, only open in the back and with half a dozen buttons on it.

Now lay the false face and the dress in the sunlight for a day or so and you will find on taking them into the dark that they shine with a ghostly radiance. This phenomenon is called *phosphorescence*, and it is caused by the luminous paint absorbing the light-waves and sending them out after the sunlight has stopped acting on it.

Next, cut out a square piece of black velvet, or canton flannel, and sew this to the top of the false face so that it will fall over the front or the back of it as you wish. Finally, make a black bag about 1 foot square and you have all the props for the materialization of *Mysteria.*

You must now look after your assistant and yourself. He must be dressed in a black canton flannel suit made like a baby's pajamas with feet in them and with a hood to match that completely covers his head, but having a couple of very small holes in it so that he can see out. He must also wear a pair of black gloves, and with this outfit on he will be quite invisible in the dark room. On the other hand, you must be dressed in either a white linen or a white flannel suit, so that you will always be visible in the aforesaid dark room.

Now just before you are ready to call this beautiful spirit from the vasty deep, fold the luminous cheese-cloth up neatly and put it into the bag, then leave it and the frame with the canton-flannel flap over the face in the outer dark room. As soon as the spectators are seated, turn out all the lights, and have your assistant bring in the wire frame and stand it silently against the wall. He then takes the luminous dress from the bag, and the audience will see it as a hazy patch of light.

As he unfolds it, the light gets brighter and larger, and as he buttons it on the wire frame it takes on the shape of the female form, but it is headless. Slowly he draws the piece of flannel up and exposes the face, and *Mysteria*, as truly a spirit of the other world as ever was materialized, appears in all her wondrous beauty and effulgent glory, as in Fig. 164.

Fig. 164.—The Spirit of Mysteria.

But we are becoming spectrally sentimental again, and this will not do, for we must get back to the hard things of this earth. Your assistant grips the spirit near the place where her feet ought to be and holds her up; then he swings her, pendulum like, from one side to the other and finally lets her come to rest in a recumbent position with her front side to the audience, of course, and there she gracefully rests until you command her to *dematerialize* and return to the place whence she came.

To perform this extraordinary feat, your assistant proceeds to take off her dress which he pins to the wall and leaves it there until he has taken the frame into the next room. Returning, he grips the dress and waves it in the air so that the audience sees the phosphorescent light high and low and everywhere at the same time. Finally, he gradually rolls the dress up and puts it under his arm when he makes his exit into the outer room. Then you turn on the lights and you will find the spectators nearly as pale but not half so beautiful as *Mysteria* herself. And thus chemistry, with a little physics thrown in, makes a spirit of a few poor "props."

CHAPTER XIV.

SAFE AND SANE FIREWORKS

Every year on the fifth of November the British celebrate *Guy Fawkes Day* with bonfires and fireworks just as we celebrate the Fourth of July, *Independence Day*, but, it is needless to say, for a wholly different reason. Guy Fawkes lived from 1570 to 1606, and he was the chief conspirator of the famous *Gunpowder Plot*, as it is called. This plot, which has ever since lived in history, was an idea that originated in the brain of one Catesby to blow up the Parliament House and in this way destroy King Charles I. On the fourth of November, which was the day set for the explosion to take place, Thomas Knyvett, a Westminster magistrate, discovered the plot and Fawkes was arrested. He was tried, together with his co-conspirators, the following January, and as he had no defense he was found guilty, and executed.

So the fifth of November is known in England as Guy Fawkes Day, and it is quite likely that our idea of celebrating Independence Day on the Fourth of July with bonfires and fireworks was taken from the old English custom that had its origin in the Gunpowder Plot. Be that as it may, you can do the following curious experiments with fire, flame, and smoke without danger if you stick to the directions, use no more of the ingredients than the formulas call for, and make them out of doors.

How to Make Fire Without a Match. Put 3 drops of glycerine ($C_3H_5(OH)_3$), *and no more*, in a pie-plate and then put 1 teaspoonful of crystals of potassium permanganate ($KMnO_4$) on top of it. In a short time the substances will react on each other, and then smoke will be evolved. If you have used the right amount of potassium permanganate ($KMnO_4$), the substances will begin to burn with a purple flame.

FIG. 165.—Writing with Fire Ink.

Writing With Fire Ink. This experiment should be made in a dark room, and is one that is quite out of the ordinary. Put a teaspoonful of water (H_2O) in a test tube, add ½ teaspoonful of potassium nitrate (KNO_3) and heat it over the flame of your alcohol lamp until the salt is dissolved. Now take a toothpick and write with the solution on a sheet of ordinary soft, porous paper, and make the lines heavy, with no break in the continuity of them.

When the paper is perfectly dry, take it in a dark room, then light a match and when it is burning well blow it out, so that only a kindling spark remains; touch the left-hand end of the writing with the match and the potassium nitrate

(KNO_3) will ignite and burn along like a fuse until the other end is reached, while the rest of the paper will not be burnt, as shown in Fig. 165.

Rapid Oxidation of Zinc. Here is another way to make a fire without a match, and this is by the rapid oxidation of zinc (Zn). Mix ½ teaspoonful of ammonium chloride (NH_4Cl) and 5 teaspoonfuls of ammonium nitrate (NH_4NO_3) on a pie-plate and then spread out the mixture in a thin layer. On top of this sprinkle 1 tablespoonful of powdered zinc (Zn) and then let a single drop of water (H_2O) fall in the center of it. The mixture will soon begin to burn, and the oxidation takes place so fast that the zinc (Zn) is ignited. It is the ammonium nitrate (NH_4NO_3) that supplies the oxygen (O) for the combustion of the zinc (Zn).

How to Make a Safe Fuse. Put a little water (H_2O) in a beaker and add as much potassium nitrate (KNO_3) to it as it will dissolve. This done, soak a soft, thick string in this solution for 10 or 15 minutes, and the salt will fill the pores of it. Now when you light one end of the string, it will burn slowly and steadily along until the other end is reached. All you need to do to make a time fuse is to use the right length of string, and this you can determine by making a trial or two.

How to Make a Flash-Light. Put ¼ teaspoonful of powdered magnesium (Mg) — no more — into the bowl of a tablespoon and hold it over the flame of your alcohol lamp, at the same time turning your face away from it; suddenly there will be a bright flash of light and in the spoon you will find a greyish powder. This substance is magnesium oxide (MgO), and it is the result of the reaction that takes place when the magnesium (Mg) combines with the oxygen (O) of the air. The experiment shows in a brilliant way the direct combination of these two elements.

How to Make Explosive Matches. For this experiment you need a few ordinary parlor matches and some sodium silicate ($Na_2Si_2O_4$), or *water-glass* as it is called. To make the latter put 1 tablespoonful each of silicon dioxide (SiO_2) or *silica*, as it is called, and sodium hydroxide ($NaOH$), that is, caustic soda, in a beaker and pour on enough boiling water (H_2O) to dissolve them; when this is done, sodium silicate ($Na_2Si_2O_4$), or water-glass is formed thus:

SiO_2	$+$	$NaOH$	$=$	$Na_2Si_2O_4$	$+$	H_2O
Silicon dioxide		Sodium hydroxide		Sodium silicate or water-glass		Water

To make the matches explosive, dip their heads in the water-glass ($Na_2Si_2O_4$), let them dry, and then dip them into melted paraffin. Now when a friend asks you for a match, hand him one of these and on striking it it will pop and sputter like a string of Liliputian firecrackers.

How to Make Rainbow Lights. Here are two very pretty experiments, but you must do them outdoors. Put ½ teaspoonful each of strontium nitrate ($Sr(NO_3)_2$), powdered charcoal (C), powdered iron (Fe), powdered magnesium (Mg), and sulphur (S), together with 1 teaspoonful of potassium nitrate (KNO_3) in a tin pan and mix them together, but *do not rub or grind them.* Now set the pan on a brick where the sparks can fly about and not do any harm. Put one end of a fuse a foot long in the mixture and light the free end of it; when the burning fuse ignites the different substances they will burn with varicolored lights and throw out brilliant scintillating sparks. The colored lights are produced by the burning metals, while the sparks are set up by the oxygen (O) liberated from the potassium nitrate (KNO_3), which oxidizes the different metals.

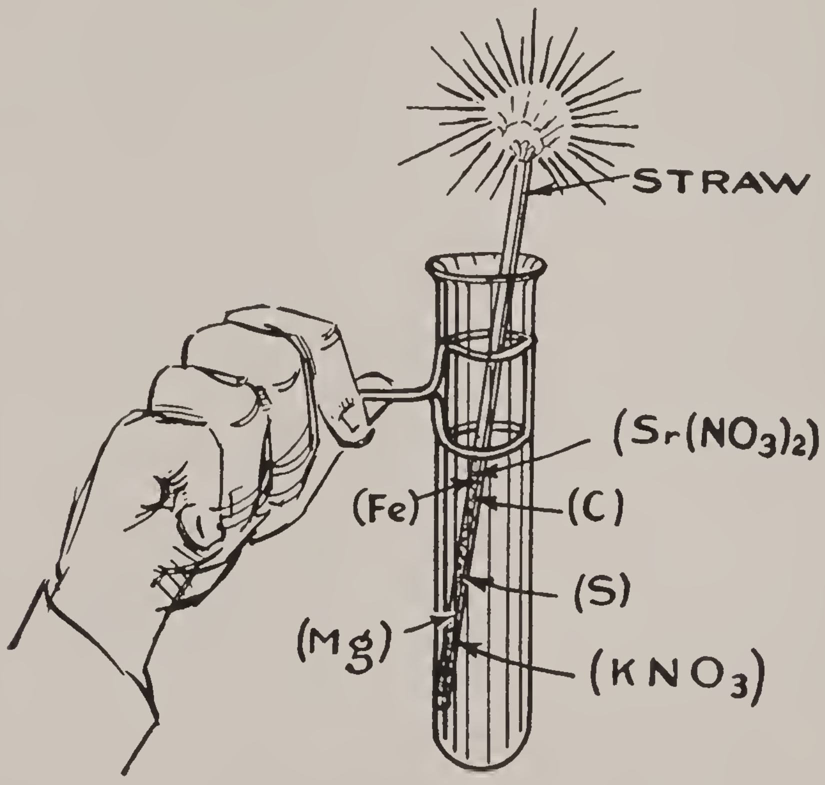

FIG. 166.—Making Rainbow Lights.

Get a paper straw, such as you use when imbibing soda-water, fold over one end, fill it two-thirds full of the mixture used in the foregoing experiment, and then set it in a test tube, as shown in Fig. 166. Now light the upper and free end of the straw, and when this mixture is ignited by it you will have a very pretty rainbow-color effect.

How to Make Fourth of July Sparklers. Make a mixture of 1 teaspoonful of potassium nitrate (KNO_3) and 2 teaspoonfuls of powdered magnesium (Mg) on a sheet of paper and stir them together, but *do not rub or grind them.* Now coat half a dozen pieces of iron wire each about 6 inches long, with melted paraffin and then roll them in turn in the mixture until they are coated all over with it. When they are cold, light the end of one of them with a match and as the magnesium (*Mg*) burns it will throw out bright scintillating sparks, as shown in Fig. 167.

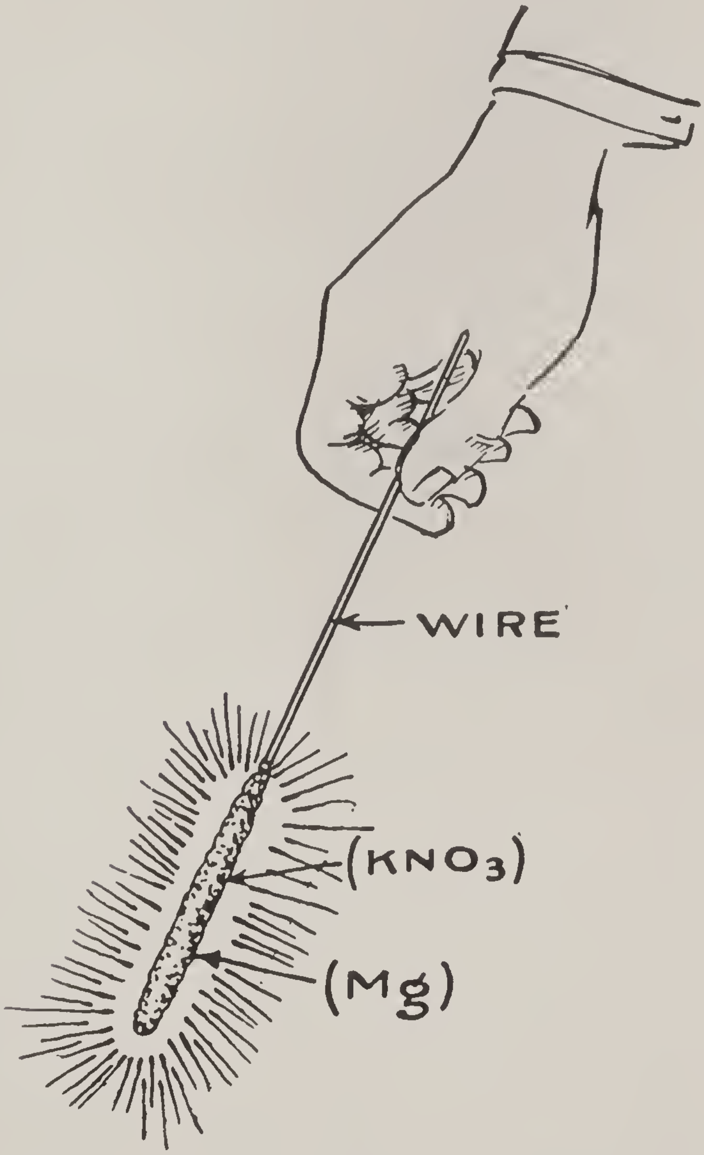

Fig. 167.—A Fourth of July Sparkler.

How to Make a White Flash-Light. Put ½ teaspoonful each of potassium nitrate (KNO_3) and powdered magnesium (Mg) into an iron frying-pan and mix, but *do not rub or grind them.* This done, sprinkle ½ teaspoonful of powdered sulphur (S) on the mixture and then light the sulphur (S) with a match fastened to a wire, as shown in Fig. 168. The burning sulphur (S) will soon ignite the potassium nitrate (KNO_3) and magnesium (Mg). They will then suddenly combine with a bright, dazzling, white flash.

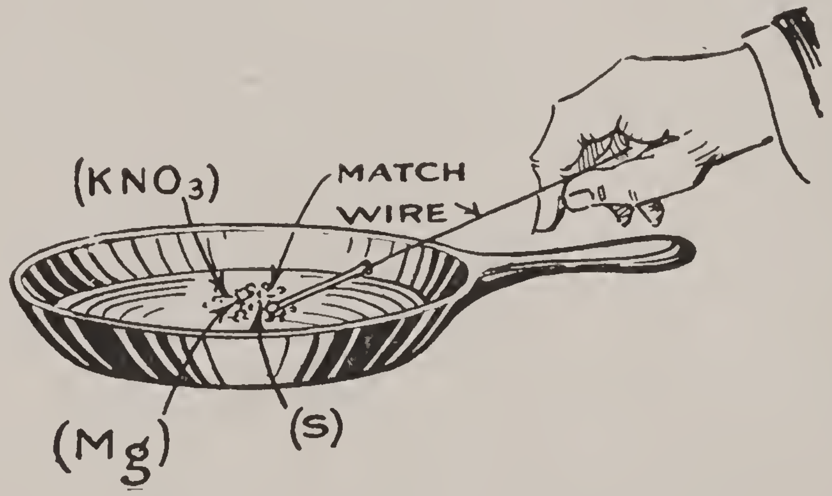

FIG. 168.—Lighting a Flash-Light.

How to Make a Red Flash-Light. To make a red flash-light use ½ teaspoonful of strontium nitrate ($Sr(NO_3)_2$) and a like amount of potassium nitrate (KNO_3). Mix them in an iron pan, taking the precautions explained in the foregoing experiment, and light the mixture. A bright, dazzling red light will be produced.

How to Make a Green Flash-Light. Mix ½ teaspoonful each of potassium nitrate (KNO_3), boric acid (H_3BO_3), or *boracic acid*, as it is commonly called, and powdered sulphur (S) and 1 teaspoonful of powdered magnesium (Mg) in an iron pan, taking the precautions explained for making a white flash-light, and fire the mixture with a fuse (see "How to Make a Fuse") and it will burn with a brilliant green flash.

How to Make Flash Paper. Magicians use this kind of paper and when they ignite it, it vanishes in a flash of light, hence, it is called *flash paper.* You can buy it for 25 cents a sheet of dealers in magical supplies, or you can make it as follows: Pour 1½ fluid ounces of sulphuric acid (H_2SO_4) and 2½ fluid ounces of nitric acid (HNO_3) into a beaker (you can measure them in a graduated glass) and stir them with a glass rod. This done, pour the solution into a 4-by-5-inch glass photographic developing-tray, and then immerse in it several sheets of thin print paper (that is, paper which is unsized and porous).

Let the sheets of paper remain in the solution for 10 minutes, then pour off the latter and wash the paper under a stream of running water (H_2O) for an hour, so that every molecule of the acids will be removed. Unless the sheets are washed thoroughly they are apt to catch fire spontaneously, and if struck with a hammer they will explode, for what you have really done is to convert the paper, which is practically pure cellulose ($C_6H_{10}O_5$) into nitro-cellulose ($C_{12}H_{14}O_4(ONO_2)_6$), and this is the gentle substance that goes by the ordinary name of *guncotton.*

After washing the sheets, drain them off and hang them on a string stretched across the room to dry. Now if you will hold a sheet by a corner with your fingers and touch the opposite diagonal corner to the flame of a candle it will disappear in a flash of light, and because the combustion is so perfect it will leave an imperceptible amount of ash behind.

How to Make Colored Flash Paper. To make colored flash paper you need only to dip the sheets after you have put them through the acid bath, and washed and dried them, in a saturated solution of the following compounds. Half fill a glass photographic tray with warm water (H_2O), and put in as much lithium chloride ($LiCl$) as it will dissolve. You will then have what is called a saturated solution; this will give the paper a *red* flash. Use a saturated solution of copper chloride ($CuCl_2$) for making *blue* flash paper. Use barium chlorate ($BaCl_2$) for *green* flash paper, and potassium nitrate (KNO_3) for *violet* flash paper.

How to Make Flash Handkerchiefs. To make a handkerchief disappear in a flash of fire, magicians use what they call a *flash handkerchief*. This is made in exactly the same way as flash paper, except that you use a handkerchief of cheese-cloth. When the latter is treated with the acid solution, it becomes a very pure form of guncotton, for cotton is a purer form of cellulose ($C_6H_{10}O_5$) than paper. As it is explosive when ignited in a confined space, do not roll it up or enclose it when you touch it off.

How to Light a Paper Without a Flame. An experiment that is a favorite with professional *fire-eaters* is to light a piece of paper by simply breathing on it. Now I submit that the breath of even a fire-eater is not nearly hot enough to raise the *kindling temperature* of a piece of paper to a point where it will catch fire. So there must be some trick in it, and here it is.

Get a glass tube that has a bore of about 1/32 inch and an outside diameter of 1/16 inch and cut it into 2-inch lengths. Seal one end of each of the tubes, as shown at

A in Fig. 169, then fill them with sulphuric acid (H_2SO_4) and seal the other end, as at B; now mix 1 teaspoonful each of sugar ($C_{12}H_{22}O_{11}$) and potassium chlorate ($KClO_3$) and put as much of this *quick-match*, as it is called, as you can get on the head of a lead pencil on a sheet of tissue paper, as at C, and then wrap the mixture and an acid tube up in it tight.

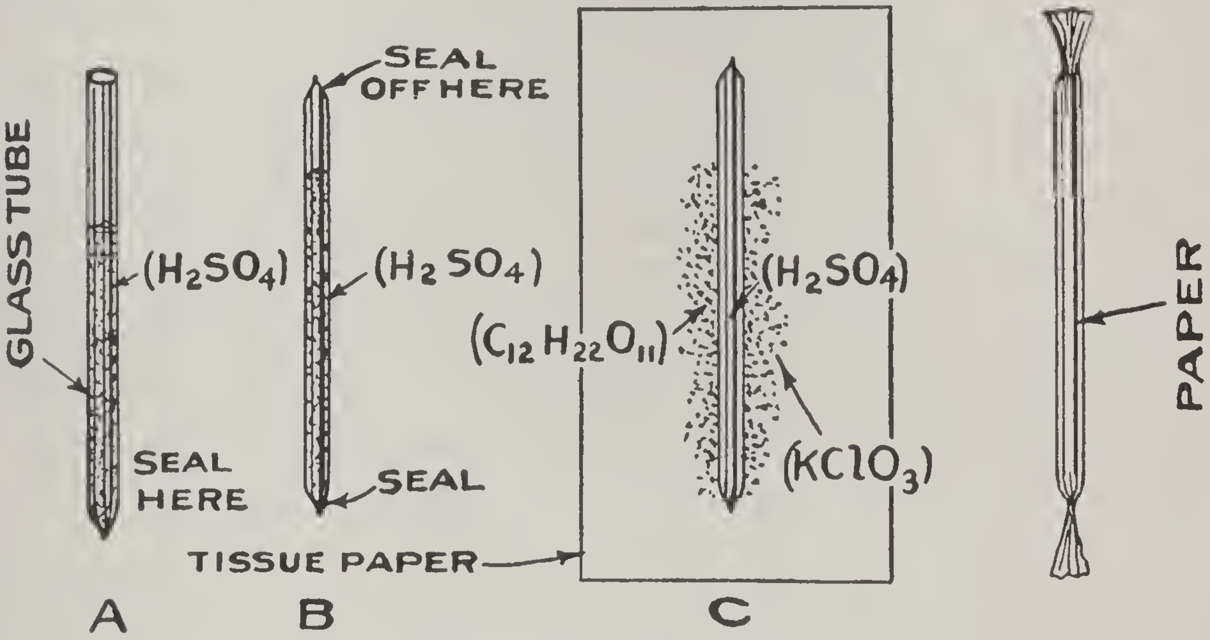

FIG. 169.—Lighting a Paper without a Match.

Now wrap this up in a sheet of ordinary paper, or one of flash paper, and when you want to ignite it with your breath you not only breathe on it (which hasn't anything to do with the case) but you squeeze the paper and break the tube. The acid in it will then come in contact with the quick-match mixture and they will instantly blaze forth and set the paper on fire. You can buy these acid tubes all ready to use of dealers in magical supplies for a very low price.

How to Light a Paper With a Piece of Ice. Fix a piece of metal potassium (K) about one-fourth the size of a pea to one corner of a perfectly dry sheet of paper by gluing a bit of paper over it, but let a corner of the metal stick out, and it cannot then be seen. When you are ready to fire the paper, press a piece of ice to the projecting corner of the potassium (K). The instant the water (H_2O) of the melting ice touches the metal, hydrogen (H) is set free and the heat of the reaction makes it catch fire. The paper will be ignited in turn.

The Great Fire-Eating Trick. When you have learned this great secret you can blow out a stream of bright sparks from your mouth "to the horror of all beholders," or at least this is what the magical catalogues say. Here is the great secret. Put 2 tablespoonfuls of water (H_2O) in a beaker and stir in as much potassium nitrate (KNO_3) as it will dissolve.

Soak a piece of thick, soft cord, about a foot long, in it over-night, then dry it thoroughly and cut it up into pieces about 1 inch long. Now light one of these pieces and roll it up loosely in a little ball of cotton about 1 inch in diameter; put this in your mouth and then blow until a shower of sparks issues forth, as shown in Fig. 170. You can heighten the effect by pretending to eat a tuft of cotton every time you blow out sparks. When you inhale, always do so through your nose, and when you exhale, always be sure to do so through your mouth, in which case the experiment will succeed beautifully.

How to Make Colored Fire.—Red Fire. Put 1 teaspoonful of strontium nitrate ($Sr(NO_3)_2$) and powdered sulphur (S), and 2 teaspoonfuls each of potassium nitrate (KNO_3) and powdered charcoal (C) in a soup-plate, or a pan, and mix them together with a stick, but *do not rub or grind them.* Make a little pile of the mixture in the center of the dish and light it with a fuse, and it will burn with a brilliant red light. It is the strontium nitrate ($Sr(NO_3)_2$) that gives the light its bright red color, while the sulphur (S) and charcoal (C) provide the material which burns, and the potassium nitrate (KNO_3) supplies the necessary oxygen (O) for them to burn in.

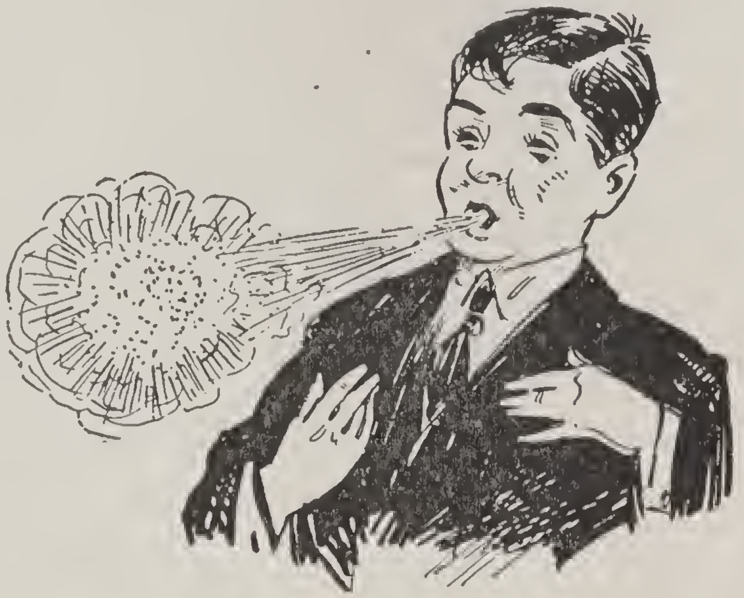

FIG. 170.—The Great Fire-Eating Trick.

Green Fire. Put 1 teaspoonful of powdered sulphur (S) and 2 teaspoonfuls each of potassium nitrate (KNO_3), powdered charcoal (C), and powdered zinc (Zn) in the dish, and mix and light them with a fuse, as in the last experiment. The mixture will then burn with a bright green color.

Yellow Fire. Put 1 teaspoonful each of sodium chloride ($NaCl$), which is common table salt, and powdered sulphur

(S) and 2 tablespoonfuls of potassium nitrate (KNO_3) and powdered charcoal (C) in a dish, or pan; mix and fire with a fuse as in the preceding experiment, and the mixture will burn with a bright yellow color.

Bengal Lights. Put 1 tablespoonful of potassium nitrate (KNO_3) 1 teaspoonful of powdered sulphur (S), and ½ teaspoonful of antimony trisulphide (Sb_2S_3) in a dish, or pan, and mix and ignite them with a fuse, as in the foregoing experiments. This mixture will then burn with a bright white light. If you will add ½ teaspoonful of powdered magnesium (Mg) to the above mixture, the light will be exceedingly brilliant.

How to Make Phosphine Smoke Rings. And I might add, a horrible odor at the same time. Put enough water (H_2O) in a glass retort to make it one-fourth full, and drop in 3 or 4 pieces of potassium hydroxide (KOH) or *caustic potash*, as it is commonly called, each of which is about as large as the stone of a cherry. Now when they are dissolved, drop in 2 bits of white phosphorus (P), each about the size of a large pea, add 1 teaspoonful of ethyl ether[1] ($C_4H_{10}O$), which is *common ether*, and put the glass stopper back in the retort.

This done, set the retort in the ring of your support-stand, have the free end of it dip into a glass dish of water (H_2O), and then set the alcohol lamp under the bowl of the retort, as shown in Fig. 171, and you are ready for the experiment. All you have to do now is to light the lamp and let the solution boil gently, and very soon a colorless gas called phosphine (PH_3), but which used to be called *phosphoreted*

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[1] When 2 molecules of alcohol lose 1 molecule of water, ether results.

hydrogen, and has an odor like that of rotten fish, will be given off; as it bubbles up through the water (H_2O) and

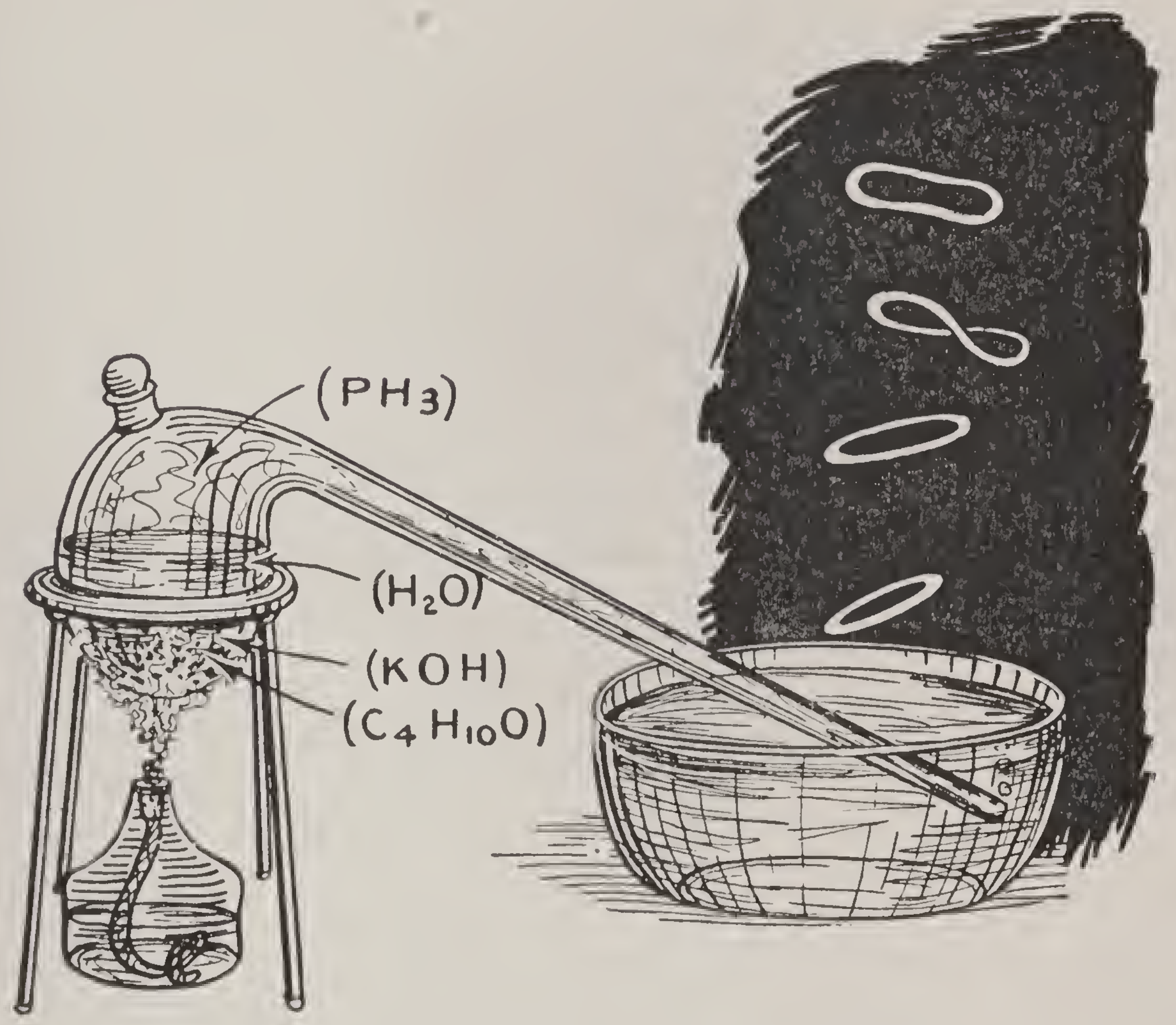

FIG. 171.—Making Phosphine Smoke Rings.

passes into the air, it will catch fire of its own accord and form wonderful rings of smoke.

The purpose of the ether ($C_4H_{10}O$) is to drive the air out of the retort and so prevent the burning of the first bubbles of gas that are formed *inside* of it. The heat must be carefully regulated, so that the bubbles will not follow each other too quickly, and the air must be perfectly still, so that the burning bubbles can form smoke rings.

How to Make Pharaoh's Serpents. This chemical reaction is a never-ending source of wonder, and it has been suggested that the ancient Egyptian conjurors in the time of Moses knew how to produce it or, at least, something very like it — hence the name *Pharaoh's Serpents*. The bald effect is that the lighting of a pill, or egg, the size of a pea, will cause a serpent-like form to wriggle forth from it with a length of several feet. This writhes about in a very lifelike manner, as shown in Fig. 172, until the egg is completely consumed.

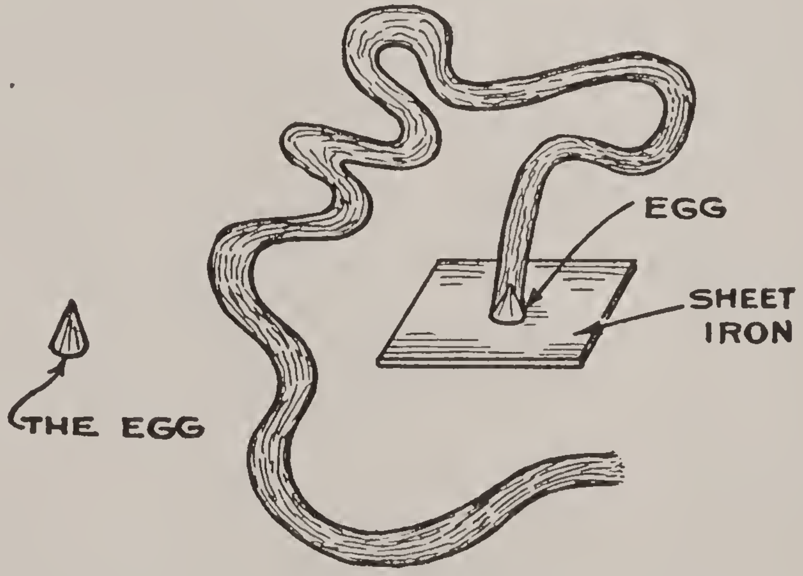

FIG. 172.—Pharaoh's Serpent Cometh Forth.

Put a drop of mercury (Hg) about the size of a pea in a test tube and add 3 or 4 drops of water (H_2O) to it, and a like amount of strong nitric acid (HNO_3). Now hold the tube over the flame of your alcohol lamp and let it heat gently; very soon the mercury (Hg) will dissolve. Then let the solution boil for ½ minute; this done, add 4 times as much water (H_2O) as there is of the solution, which now contains mercuric nitrate ($Hg(NO_3)_2, H_2O$).

Next, dissolve as much potassium thiocyanate ($KNCS$)

as you can get on the head of a lead pencil in ½ teaspoonful of water (H_2O) and add this to the solution of mercuric nitrate ($(HgNO_3)_2, H_2O$). When this is done, a dirty grey precipitate will be thrown down, but as you keep adding more of the potassium thiocyanate ($KNCS$) to the mercuric nitrate ($(HgNO_3)_2, H_2O$) solution, the precipitate will become a creamy white.

Now filter the solution and save only the precipitate which remains behind on the filter paper, then wash it by pouring on a little water (H_2O), and let it filter through. Let the latter drain off, then take the paper out of the funnel, unfold it, lay it on a couple of sheets of blotting paper and let it stay there until the precipitate is perfectly dry. scrape the latter off into a small dish and put a drop or two of mucilage on it so that you can mold it with your fingers into little cone-shaped pieces about as large as peas.

Finally, let them dry thoroughly, and you have the eggs of the famous Pharaoh's serpents. Now light one of them with a match, and it will burn with a nearly invisible flame, and at the same time form a brown ash of almost incredible length which curls up and twists round after the manner of a live serpent, whence it gets its name.

NOTE.— As potassium thiocyanate ($KNCS$) is a poison, be sure to wash your hands after you have shaped the compound into eggs. Further, as it gives off poisonous gases when it burns, do not get close enough to inhale them. Either make the experiment out of doors, or in a fireplace.

Here is a way to make Pharaoh's serpents which are not poisonous, neither are they anywhere nearly so effective as the foregoing. Mix, but *do not rub or grind*, ½ teaspoonful each of powdered potassium nitrate (KNO_3) and sugar ($C_{12}H_{22}O_{11}$) and 1 teaspoonful of powdered potassium dichromate (K_2CrO_7). Now add just enough mucilage to make a paste of the mixture and then shape them into little cones. When you light these, or the kind described above, always do so at the top of the cone.

CHAPTER XV

USEFUL HOUSEHOLD RECIPES

In this chapter it is my intention to tell you how to make some interesting experiments that have to do with things in and around the house, and the family living in it. These experiments include the making of soap, water-softeners, cleansing compounds, disinfectants, dyes and inks, together with a number of miscellaneous recipes.

How to Make Soaps. In Chapter VIII I told you how to make hard and soft soaps simply as experiments in chemistry, and here I shall give you some additional easy formulas for making other kinds of soap, but also on a very small scale.

Toilet Soap. Put a tablespoonful of olive oil ($C_3H_5(CO_2C_{17}H_{33})_3$)[1] into a small porcelain evaporating-dish and then pour the same amount of alcohol (CH_4O) over it; next, put a teaspoonful of sodium hydroxide ($NaOH$), that is, caustic soda, in a test tube and pour a like amount of water (H_2O) over it. Now put 20 drops of this solution in the dish with the other two compounds.

This done, heat the dish gently until the solution boils and all the alcohol (CH_4O) has evaporated, which you will know when you can no longer smell the odor from it; evaporate the solution slowly until the remaining mass is quite

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[1] This is the formula for *olein*, and olive oil contains 75 per cent of it.

dry, and this is, or at least it should be, *soap*. If it has not *saponified*, that is, changed into soap, put a little more alcohol (CH_4O) and sodium hydroxide ($NaOH$) in the dish and boil it again.

Perfumed Soap. Take a piece of good hard soap the size of a walnut and melt it in a test tube, or the tin cover of a baking-powder can, and while it is in a liquid state add a few drops of perfume of any kind to it and stir it in thoroughly. When it is cold you can easily detect the odor by smelling of it, and it will be very much in evidence when you wash with it.

Colored Soap. Follow the same directions as for making the perfumed soap given above, except that you add a harmless coloring matter to it while it is in a melted state. To give the soap a red color, put ¼ teaspoonful of cochineal[1] in a test tube one-fourth full of water (H_2O) and boil it until the solution is a bright red. Now put enough of this into the melted soap to give it the tint you want. Other colors can be had by using vegetable dyes of various kinds.

Floating Soap. Put enough good hard soap into a test tube to fill it half full and then melt it. Stick a straw or glass tube into the solution and just before it gets hard, blow a blast of air through it and stir it at the same time; this will fill it with air bubbles, and it will then be lighter than water (H_2O) and, consequently, the soap will float.

Glycerine Soap. Cut up a lump of good soap the size of a walnut and put it in a test tube and melt it. Then add ½ teaspoonful of glycerine ($C_3H_5(OH)_3$) and stir them

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[1] This is a brilliant scarlet dye stuff made by killing female cochineal insects and drying them.

until they are thoroughly mixed. When this is cold you will have glycerine soap.

Sapolio. Cut up a piece of soap the size of a walnut and melt it; then add 5 or 6 times the amount of very fine sand (SiO_2),[1] together with a bit of glue, and mix them thoroughly; while the mixture is still hot put it into a little mould made of wood, or a tin box will do, lay a piece of wood on top of it and set a flat-iron or other weight on that. This done, let it dry thoroughly and you will have a cake of *sapolio*, or a close approximation to it.

How to Make a Safe Dry-Cleansing Compound. The process of cleaning goods with solvents other than water (H_2O) is called *dry cleansing*. Gasoline (C_7H_{16}) and benzine (C_8H_{18}), which are *hydrocarbons* obtained from *petroleum*, are very good solvents for oil, grease, tar, and other like organic matter, and they are largely used for removing them from clothing, but they are dangerous because they are easily ignited and explosive.

You can make a cleansing solution which will not burn, by adding 1 ounce of carbon tetrachloride (CCl_4), which is a liquid compound made by passing dry chlorine (Cl) into carbon disulphide (CS_2), to 5 ounces of benzine (C_8H_{18}). Or you can use carbon tetrachloride (CCl_4) alone, for while it is not quite as cheap, it is even safer and it evaporates about as quickly.

How to Take Out Spots and Stains.—A Fresh Grease Spot. Lay a piece of blotting paper over the grease spot and press on it with a hot flat-iron; the heat will melt the

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[1] This is the formula for *silicon dioxide*, or *silica*, as it is called, and *sand* is composed chiefly of it.

grease and the blotting paper will absorb it. Hence this is not a chemical experiment but one that has to do with physics. As long as the spot is gone, it really doesn't matter.

Old Grease Spots. You can remove an old grease spot from clothing by dissolving it out with alcohol (CH_4O), benzine (C_8H_{18}), carbon tetrachloride (CCl_4), or the solution described above. In taking out a grease spot, start at the edge of it with the cloth saturated with the solvent, and then keep on working toward the center of it.

Paint Spots. The first thing to do is to soften the paint, and this can be done by pouring on a little carbon tetrachloride (CCl_4); after it has soaked for a while, moisten a bit of clean muslin with turpentine ($C_{10}H_{16}$) and rub the spot until all traces of the paint have disappeared.

Ink Spots. To take out an ink spot on woolen clothing, rub it lightly with a bleaching solution made by dissolving 1 teaspoonful of calcium hypochlorite ($Ca(OCl)_2$), that is *chloride of lime*, in 2 tablespoonfuls of water (H_2O). This will bleach out the black spot and leave a yellow spot, and this you can remove by soaking a pellet of cotton in hydrogen dioxide (H_2O_2) and with it gently rubbing the spot, which in turn will disappear.

Where fresh ink is spilled on bright-colored goods, or on a carpet, it can generally be removed by repeatedly washing the stain with fresh, sweet milk.[1]

To remove ink from paper, dissolve ½ teaspoonful each of tartaric acid ($C_4H_6O_6$) and calcium hypochlorite (Ca

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[1] Milk is an emulsion formed of 80 to 90 per cent of water in which there is dissolved 2 to 6 per cent of *casein*, ⅕ to 9 per cent of *milk-sugar*, 1 to 2 per cent of *mineral salts*, and 2½ to 6 per cent of *fat;* and it swarms with bacteria.

$(OCl)_2$), in 2 tablespoonfuls of water (H_2O). Now take a pointed glass rod, or a wood toothpick will do, dip it into the solution and rub with it the ink that you want to remove, and it will fade away. The tartaric acid ($C_4H_6O_6$) and the calcium hypochlorite ($Ca(OCl)_2$) react on each other and set the chlorine (Cl) free. This with the water (H_2O) makes hypochlorous acid ($HClO$), which, as you know, is a bleaching agent. There are some kinds of ink that cannot be bleached out with this solution.

Iron-Rust Stains. Rub the stain with a solution made of 1 teaspoonful of oxalic acid ($C_2H_2O_4$) dissolved in 3 tablespoonfuls of water (H_2O). When the stain has been removed, wash out the acid solution with a plentiful supply of water (H_2O).

Alkali Spots. Where an alkali, such as sodium hydroxide ($NaOH$), that is, caustic soda gets on a piece of goods you can take it out by rubbing it gently with a piece of clean muslin dipped in the oxalic-acid solution, made as described above. After the acid has neutralized the alkali, causing the spot to disappear, wash it out with plenty of water (H_2O).

Mildew Stains. You can remove mildew stains by rubbing them gently with a solution made by dissolving 1 teaspoonful of calcium hypochlorite ($Ca(OCl)_2$) in a test tube half full of water (H_2O). It will then bleach out the stains; after which the goods should be washed in a plentiful supply of water (H_2O).

HOW TO MAKE BLEACHING COMPOUNDS.

For Cotton and Linen Goods. Hypochlorous acid ($HClO$) is the universal bleaching compound for cotton and linen goods. You can make it by dissolving a teaspoonful of calcium hypochlorite ($Ca(OCl)_2$), which is chloride of lime, or bleaching powder, as it is called when used for this purpose, in ½ pint of water (H_2O).

For Wool and Silk. Never try to bleach wool or silk with bleaching powder, or any compound that makes hypochlorous acid ($HClO$), for this destroys these kinds of goods because they contain organic compounds called *protein* ($CHOHS$). To bleach wool and silk, use sulphurous acid (H_2SO_3), which you must not confound with sulphuric acid (H_2SO_4). Sulphurous acid (H_2SO_3) is formed by dissolving sulphur dioxide (SO_2) in water (H_2O) thus:

$$\underset{\text{Sulphur dioxide}}{SO_2} + \underset{\text{Water}}{H_2O} = \underset{\text{Sulphurous acid}}{H_2SO_3}$$

While sulphurous acid (H_2SO_3) is like sulphuric acid (H_2SO_4) except that it contains one less molecule of oxygen (O), it differs from it in that it is a very weak acid. It bleaches by virtue of the fact that it combines with various coloring substances and makes other compounds, which process leaves the goods white.

For Hair and Wool. For bleaching hair and wool, use hydrogen peroxide (H_2O_2), which, as its formula shows, is very like water (H_2O), except that it has 2 molecules of oxygen (O) where the latter has only 1 of oxygen (O). This difference is enough to make it heavier than water (H_2O), give it a syrupy consistency, and it makes hydrogen peroxide (H_2O_2) a powerful bleaching compound. It is made by treating barium dioxide (BaO_2) with sulphuric acid (H_2SO_4).

How to Make Disinfectants. A disinfectant is a substance that will kill the germs which cause various diseases. Among the better-known disinfectants are chlorine (Cl), sulphur (S), hydrogen peroxide (H_2O_2), formaldehyde (CH_2O) and phenol (C_6H_5OH), or *carbolic acid*, as it is popularly called.

Chloride of lime ($Ca(OCl)_2$) is a good disinfectant, and you need only to dissolve 1 ounce of it in 1 quart of water (H_2O) to make it. Where there are germs of malignant diseases, the rooms can be disinfected by burning sulphur (S) in them. The sulphur (S) combines with the oxygen (O) of the air, and this forms sulphur dioxide (SO_2). To make this gas effective you must seal up the windows and doors of the room you want to disinfect by pasting strips of paper over the cracks, then put a couple of lumps of sulphur (S), about the size of walnuts, in an iron pot and ignite them.

Hydrogen peroxide (H_2O_2) is not only a bleaching agent but it has the remarkable property of destroying tissues of the body that are dead or decaying, while it will not affect healthy, living tissues. Another good feature about it is that when it reacts on dead and decaying tissue, water (H_2O) only is left behind, and, hence, there is nothing which will irritate or poison the tissues that are living. For this reason, it is very much superior to disinfectants of other kinds. Use a 3-per-cent solution of hydrogen peroxide (H_2O_2) for disinfecting wounds and sores, and this you can get at any drug store.

While formalin ($CH_2O + H_2O$), which is a solution made by dissolving 40 per cent of formaldehyde (CH_2O), a gas, in 60 per cent of water (H_2O), is often used as a preservative of milk, it is harmful when taken into the system, but it is a very good disinfectant. You can get formalin ($CH_2O + H_2O$) all ready to use at the drug store.

Finally, phenol (C_6H_5OH), or carbolic acid, to give it its common name, is a most excellent disinfectant. It is one of the products of coal-tar, and to make a disinfectant of it you need only to mix 5 per cent of it with 95 per cent of water (H_2O). In use this disinfectant is sprinkled around.

How to Make and Use Dyes. Dyes are of two general kinds: *natural colors* and *artificial colors*. The former are made of various plant, animal, and mineral matter, while the latter are either extracted from *coal-tar* or else they are made *synthetically*.

Logwood (*red*), indigo[1] (*blue*), and tumeric (*yellow*) are some of the plant colors; cochineal (*scarlet*) is a dye made from insects; chrome (*green* and *yellow*), iron buff, prussian blue, and manganese brown are mineral colors. *Aniline dyes* are made from coal-tar, and other dyes, as for instance indigo, are made synthetically; that is, the chemist builds up a compound exactly like the one that nature makes by combining the same elements of which it is formed.

HOW TO MAKE AND USE NATURAL-COLOR DYES.

Direct, or Substansive Dyes. Nearly all the plant and animal dyes can be made by boiling the dye-stuffs in water (H_2O). The goods to be dyed are immersed in these, and

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[1] Indigo was formerly obtained from the *indigo plant*, which was extensively grown in India and Egypt, but practically all that is used now is made synthetically.

hence they are called *direct*, or *substantive*, dyes. Very often the color can be changed or improved by adding some other compound, thus:

Red Logwood Dye. Put ½ teaspoonful of logwood in a test tube half full of water (H_2O) and boil it for several minutes. This done, put ¼ teaspoonful of cobalt chloride ($CoCl_2, H_2O$) in another test tube one-third full of water (H_2O); now pour the first solution into the second one, and you will have a dark-red dye.

Black Logwood Dye. For this dye, use the same amount of ferric ammonium sulphate ($(NH_4)_2SO_4, Fe_2(SO_4)_3, H_2O$) instead of cobalt chloride ($CoCl_2, H_2O$) given above.

Green Logwood Dye. For this dye use copper sulphate ($CuSO_4, H_2O$) instead of cobalt chloride ($CoCl_2, H_2O$), as given for the red logwood dye.

Yellow Tumeric Dye. Put a teaspoonful of *tumeric* in a test tube half full of water (H_2O). Boil it for several minutes, and to this solution add a few drops of acetic acid (HCO_2CH_3).

Brown Tumeric Dye. Add ¼ teaspoonful of sodium carbonate (Na_2CO_3, H_2O) to the above tumeric solution.

NOTE.— When an acid is added to a tumeric solution, it makes a yellow dye, and when an alkali is added to it, it makes a brown dye.

Bright-Red Cochineal Dye. Put ¼ teaspoonful of *cochineal* in a test tube half full of water (H_2O) and boil it for a few minutes.

Orange Cochineal Dye. Add ⅓ teaspoonful of tartaric acid ($C_4H_6O_6$) to the foregoing dye and shake the test tube thoroughly.

Violet Cochineal Dye. Add ⅓ teaspoonful of sodium carbonate (Na_2CO_3, H_2O) to the red cochineal dye, and shake the test tube well.

NOTE.— From this you will see that when an *acid* is added to a cochineal solution, it turns it an *orange* color, and when an *alkali* is added, it turns the solution to a *violet* color.

Insoluble Dyes. Different from the above natural colors, indigo blue ($C_{16}H_{10}N_2O_2$) and *chrome*, which latter is a metallic color, *will not* dissolve in water (H_2O) and, hence, these, and others like them, are called *insoluble dyes*. But indigo white ($C_{16}H_{12}N_2O_2$) will dissolve in water (H_2O).

To Dye Indigo Blue. Dissolve ½ teaspoonful of indigo white ($C_{16}H_{12}N_2O_2$) in a test tube of hot water (H_2O) and then soak a strip of muslin in it. Take it out and hang it up in the air to dry, and the oxygen (O) of the latter will oxidize it. This changes it into indigo blue ($C_{16}H_{10}N_2O_2$), which is a fast color.

To Dye Tumeric Yellow. To dye a strip of muslin a beautiful permanent yellow put ¼ teaspoonful of lead acetate ($Pb(CO_2CH_3)_2, H_2O$) in a test tube nearly full of water (H_2O) and boil the cloth in this solution for a few minutes; this done, put ¼ teaspoonful of potassium chromate (K_2CrO_4) in another test tube nearly full of water (H_2O) and heat it. While the solution is boiling-hot, put the strip of goods in it and let it soak for a few minutes.

When the potassium chromate (K_2CrO_4) comes in contact with the lead acetate ($Pb(CO_2CH_3)_2, H_2O$) in the goods they react on each other, and yellow lead chromate ($PbCrO_4$) is formed; this latter compound is a precipitate and fills the fibers of the goods with a yellow color. As the lead chromate ($PbCrO_4$) will not dissolve in water (H_2O), it cannot be washed out, and, so the color is a fast one.

Mordant, or Adjective Dyes. Again different from insoluble dyes are the *mordant*, or *adjective* dyes, as they are called. The word *mordant* comes from a Latin word which means *to bite*, and it is a substance that fixes, or bites a color in the goods. Take three test tubes, each of which is nearly full of water (H_2O). In the first, dissolve as much aluminum sulphate ($Al_2(SO_4)_3, H_2O$) as you can get on the head of a lead pencil. In the second, dissolve the same amount of ferric chloride ($FeCl_3$). In the third, dissolve the same amount chromic acetate ($Cr(CO_2CH_3)_3$). You can hasten the action by heating the solutions.

Now dissolve ¼ teaspoonful of alizarin[1] ($C_{14}H_8O_4$), commonly called *madder*, and which is an orange-yellow dye, in each of the tubes; you will now have in the first one a *red* dye known as *Turkey red*, in the second a *violet* dye, and in the third one a *maroon* dye. This done, immerse a strip of goods in each of these different solutions, and the coloring matter will be absorbed by the mordant, and together they form an insoluble dye called a *lake* in the fibers, and so each strip is dyed a permanent color.

How to Make and Use Aniline Dyes. The simplest way to make dyes is to use *aniline colors*, and these are products of coal-tar. You can usually get these colors at a drug store, but if you actually want to dye a garment, the best

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[1] Alizarin is the active coloring matter of *madder*, and 50 years ago this plant was largely grown and used as a dye to produce the well-known color called *Turkey red*. Alizarin is now made from *anthracene*, which is a coal-tar product.

way to go about it is to buy dyes already put up in packets.

Direct Aniline Dyes for Cotton Goods. For experimental purposes, take whatever color of aniline dye you want, say, black, red, green, blue, or yellow that will dissolve in water (H_2O), and to get this kind you must ask for *direct aniline dyes*. Now nearly fill a test tube with water (H_2O) and heat it until it boils; then add a few grains of the aniline dye at a time until you have produced the depth of color you want. This done, dip a strip of muslin or other cotton goods in the dye while it is still very hot, and the job is done.

Mordant Aniline Dyes for Cotton Goods. To dye cotton goods with aniline dyes and fix them with a mordant, you must ask for *basic aniline dyes*. Put ¼ teaspoonful of tannic acid ($C_{14}H_{10}O_9$) in a test tube full of boiling water (H_2O), and then put a few grains of basic aniline dye in another test tube half full of water (H_2O) to give you the color you want. Now dip a strip of muslin in the mordant and let it soak for 5 minutes or so, then take it out and dip it in the aniline dye, and the color will be fixed there.

Acid Colors for Silk and Woolen Goods. To dye silk and wool you must get aniline dyes that are sold under the name of *acid colors*. Add a few grains of the dye to a test tube nearly full of boiling-hot water (H_2O) until you have the desired color, and then add as much sodium chloride ($NaCl$), that is common table salt, as you can get on the head of a lead pencil. Now dip a strip of silk or woolen goods in the dye while it is hot, take the strip out and dip it in a solution of tannic acid ($C_{14}H_{10}O_9$), and the dye will be fixed there.

How to Make Inks.— Black Ink. Put ½ teaspoonful of tannic acid ($C_{14}H_{10}O_9$) in a test tube two-thirds full of water (H_2O); then put ½ teaspoonful of ferric ammonium sulphate ($(NH_4)_2SO_4, Fe_2(SO_4)_3, H_2O$) in another test tube two-thirds full of water, add ¼ teaspoonful of *gum arabic* to it, and heat the contents to dissolve them.

This done, pour the two solutions into a beaker and stir them well with a glass rod, and then add a couple of drops of *oil of wintergreen* to keep the ink from spoiling. Fill a bottle with the solution and you will have a good black ink. The moment the solutions come in contact, they react on each other and form ferric tannate ($Fe(SO_4)_3$), or *iron tannate,* as it is called, and it is this compound that makes the ink black.

Blue Ink. To make blue ink, dissolve ½ teaspoonful ferric ammonium sulphate ($(NH_4)_2SO_4, Fe_2(SO_4)_3, H_2O$) in a test tube half full of water (H_2O) and then dissolve ½ teaspoonful of sodium ferrocyanide ($Na_4Fe(CN)_6, H_2O$) in a test tube half full of water (H_2O). This done, pour one solution into the other, and the reaction set up will form a blue precipitate, which is ferro-ferricyanide.

Purple Ink. Put 1 teaspoonful of *logwood* into a test tube two-thirds full of water (H_2O) and boil it until the coloring matter is well out of it; now add ½ teaspoonful of aluminum sulphate ($Al_2(SO_4)_3$) and boil it again, and a fine purple ink will result. In this ink, a *lake* is made by the combination of the *plant matter*, that is, the logwood, with a *metal,* that is, with the aluminum sulphate ($Al_2(SO_4)_3$).

Red Ink. Make the purple ink just described and then add 1 teaspoonful of sodium bisulphate ($NaHSO_4$) to it, and you will have a red ink.

Green Ink. Put ½ teaspoonful of nickel ammonium sulphate ($(NH_4)_2SO_4, NiSO_4, H_2O$) and ¼ teaspoonful of sodium ferrocyanide ($Na_4Fe(CN)_6, H_2O$) in a test tube half full of water (H_2O) and shake it until they are thoroughly dissolved. This done, put in ⅛ teaspoonful of ferric ammonium sulphate ($(NH_4)_2SO_4, Fe_2(SO_4)_3, H_2O$), and again shake it until this is dissolved, and you will have a beautiful green ink.

NOTE.— In writing with any of the above inks, always use a perfectly clean pen.

Printer's Ink. Put 1 teaspoonful of sodium silicate (Na_2SiO_3), or water-glass, as it is called and ½ teaspoonful of lampblack (C), which is the soot formed by burning oil residues, in your mortar and rub them together with the pestle until they are thoroughly mixed; put this mixture in a test tube and then fill it half full of water.

This done, stir in ¼ teaspoonful each of ferric ammonium sulphate ($(NH_4)_2SO_4, Fe_2(SO_4)_3, H_2O$), and tannic acid ($C_{14}H_{10}O_9$), and then shake the tube vigorously until a thick black liquid is formed; finally pour it out on your evaporating-dish and let it remain exposed to the air until the ink is of the proper consistency.

SOME OTHER USEFUL RECIPES.

How to Make a Liquid Ink Eraser. There are two ways to erase writing done with ink, and these are with a *rubber* or *steel ink eraser*, and with a *liquid bleaching compound.* The latter usually makes the cleaner job when it is properly done. To make a liquid ink eraser, put ¼ teaspoonful each of tartaric acid ($C_4H_6O_6$) and calcium hypochlorite ($Ca(OCl)_2$), that is, bleaching powder, in a test tube one-third full of water (H_2O) and shake it well to dissolve them.

Now take a camel's-hair brush, dip it into the solution and wash it over the writing that you want to remove. It will quickly disappear, leaving no trace. The reaction that takes place is this: the calcium hypochlorite ($Ca(OCl)_2$) and the tartaric acid ($C_4H_6O_6$) combine, and in so doing chlorine gas (Cl) is set free. This forms hypochlorous acid ($HClO$) when it comes in contact with the water (H_2O) that is in the pores of the paper.

How to Make a Good China Cement. To make a cement for mending broken chinaware, take ½ teaspoonful of *albumen*, that is, the white of an egg, and 1½ teaspoonfuls of calcium carbonate ($CaCO_3$) and mix them thoroughly together. To cement two or more pieces of chinaware together, clean the broken edges with hot water (H_2O) and let them dry; now coat the edges with the cement, press the pieces together, and then let them dry for 48 hours, and a very firm joint will be made.

How to Make an Adhesive Paste. Put 3 teaspoonfuls of powdered starch ($C_6H_{10}O_5$) in a test tube one-third full of water (H_2O) and stir to a smooth paste. Now add ⅓ teaspoonful of calcium chloride ($CaCl_2$) to a test tube one-third full of boiling water (H_2O). Next, pour this latter solution into the first test tube and then bring it to a boil, add a drop or two of oil of wintergreen, to keep it sweet, and pour it into a bottle.

How to Make Fire-Extinguishing Compounds. Among the chief fire-extinguishing compounds are water (H_2O), carbon dioxide (CO_2), and carbon tetrachloride (CCl_4). Now make the following experiment: Light a sheet of paper and place it on an old plate, then let some water (H_2O) trickle on it, and you will see that the blaze rapidly goes out. This is because water (H_2O) absorbs a considerable amount of the heat, keeps the temperature below the kindling point, and the steam (H_2O) that is formed prevents the air from supplying more oxygen (O) to it.

Now light another piece of paper and direct a jet of carbon dioxide (CO_2) on it, and the flame will be quickly extinguished. This is because the carbon dioxide (CO_2) will not support combustion and is heavier than the air; hence it soon forms a blanket over the fire, and as this prevents the oxygen (O) from feeding the flames, the latter cannot burn.

In the usual kind of *hand fire-extinguisher*, the can to which the nozzle is connected is filled with a weak solution of sodium carbonate (Na_2CO_3, H_2O), and in the top of the can there is a bottle filled with sulphuric acid (H_2SO_4); now, when you turn the tank upside down the acid runs into the sodium carbonate solution and this sets free the carbon dioxide (CO_2) that is in the latter. As the gas is generated in large quantities it develops a high pressure, and at the same time some of it is dissolved in the water (H_2O), so that both of them are forced out in a stream and put out the fire, as explained above.

Light a sheet of paper and then let a little stream of carbon tetrachloride (CCl_4) play on it, which you can do with a pipette. The instant this compound comes in contact with the flames it will put them out. This is because it forms a blanket of unburnable gas around the flames and so shuts out the oxygen (O). A new kind of hand fire-extinguisher that uses carbon tetrachloride (CCl_4) for the liquid is now on the market.

How to Clean Silverware Chemically. Put 1 teaspoonful of sodium thiosulphate[1] ($Na_2S_2O_3, H_2O$) in a test tube nearly full of water (H_2O) and shake it well. Now moisten a clean piece of cloth with the compound and with it rub the silverware to be cleaned. The film on it, which is silver sulphide (Ag_2S), will react with the sodium thiosulphate ($Na_2S_2O_3, H_2O$), causing thiosulphuric acid ($H_2S_2O_3$) to be set free and the sulphur (S) to be precipitated, both of which are easily wiped off, leaving the silver clean and bright.

How to Clean Silverware Electrically. Put the silver article to be cleaned in a zinc (Zn) or an aluminum (Al) pan or kettle and pour on enough water (H_2O) to cover it; this done, add 1 teaspoonful of sodium chloride ($NaCl$), which is common salt, for each pint of water (H_2O). This done, let the water (H_2O) in the pan or kettle boil for a couple of minutes, then take out the silver article and wash it in clean water (H_2O). It will then be as clean and bright as new. In this process, the reaction is an *electrolytic* one, that is, it is done by the action of an electric current, the pan or kettle serving as the negative pole, the silver article as the positive pole, and the salt solution as the battery solution, or *electrolyte*, as it is called. The film of

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[1] This is the so-called hypo that is used for fixing negatives in photography. See Page 222.

silver sulphide (Ag_2S) on the article is removed by *electrolysis*, that is, it is deposited on the zinc (Zn) or aluminum (Al) vessel by electrolytic action.

How to Waterproof Goods. Make a solution by dissolving 1 teaspoonful of aluminum acetate ($Al(CO_2CH_3)_3$) in a test tube half full of water (H_2O), and then soak a strip of muslin in it. This done, hold the muslin for a few minutes over the spout of a teakettle from which live steam (H_2O) is issuing. The steam (H_2O) plus the aluminum acetate ($Al(CO_2CH_3)_3$) combines and forms aluminum hydroxide ($Al(OH)_3$), which is precipitated, and this fills the hollow fibers of the cotton and makes them non-absorbent to such an extent that water (H_2O) has little or no effect on the goods.

How to Fireproof Goods. All you need to do to make a piece of goods unburnable is to soak it in sodium silicate ($NaSiO_3$), or water-glass, as this compound is commonly called. If you want to fireproof a board, use a paint-brush and coat it with water-glass. The way to make water-glass is explained in Chapter XIV.

How to Make a Hair-Remover. Druggists sell *depilatory compounds* of various kinds for removing superfluous hair, but you can make one easily and cheaply that is guaranteed harmless. Get ⅓ ounce of calcium sulphide (CaS) at the drug store and mix it with enough water (H_2O) to make a thick paste. Now spread this compound on that part of your face where the offending hair is and leave it there over-night. In the morning, the hair that you want to get rid of will be gone, and, your face will not be injured.

In this operation, the water (H_2O) and calcium sulphide

(CaS) react on each other and form calcium hydrosulphide ($Ca(OH)S$) and calcium hydroxide ($Ca(OH)_2$), that is, slaked lime. Now since both alkalis and hydroxides will decompose *proteins*, and as hair is made up of this substance, it is removed by them.

LIST OF ELEMENTS AND THEIR SYMBOLS.

Aluminum	*Al*
Antimony	*Sb*
Argon	*A*
Arsenic	*As*
Barium	*Ba*
Bismuth	*Bi*
Boron	*B*
Bromine	*Br*
Cadmium	*Cd*
Caesium	*Cs*
Calcium	*Ca*
Carbon	*C*
Cerium	*Ce*
Chlorine	*Cl*
Chromium	*Cr*
Cobalt	*Co*
Columbium	*Cb*
Copper	*Cu*
Dysprosium	*Dy*
Erbium	*Er*
Europium	*Eu*
Fluorine	*Fl*
Gadolinium	*Gd*
Gallium	*Ga*
Germanium	*Ge*
Glucinum	*Gl*
Gold	*Au*
Helium	*He*
Holmium	*Ho*
Hydrogen	*H*
Indium	*In*
Iodine	*I*
Iridium	*Ir*
Iron	*Fe*
Krypton	*Kr*
Lanthanum	*La*
Lead	*Pb*
Lithium	*Li*
Lutecium	*Lu*
Magnesium	*Mg*
Manganese	*Mn*
Mercury	*Hg*
Molybdenum	*Mo*
Neodymium	*Nd*
Neon	*Ne*
Nickel	*Ni*
Niton (radium emanation)	*Nt*
Nitrogen	*N*
Osmium	*Os*
Oxygen	*O*
Palladium	*Pd*
Phosphorus	*P*
Platinum	*Pt*
Potassium	*K*
Praseodymium	*Pr*
Radium	*Ra*
Rhodium	*Rh*
Rubidium	*Rb*
Ruthenium	*Ru*
Samarium	*Sm*
Scandium	*Sc*
Selenium	*Se*
Silicon	*Si*
Silver	*Ag*
Sodium	*Na*
Strontium	*Sr*
Sulphur	*S*
Tantalum	*Ta*
Tellurium	*Te*
Terbium	*Tr*
Thallium	*Tl*
Thorium	*Th*
Thulium	*Tm*
Tin	*Sn*
Titanium	*Ti*
Tungsten	*W*
Uranium	*Ur*
Vanadium	*V*
Xenon	*Xe*
Ytterbium	*Yb*
Yttrium	*Y*
Zinc	*Zn*
Zirconium	*Zr*

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