THE BOYS' BOOK OF CHEMISTRY

CHARLES RAMSAY CLARKE

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THE BOYS' BOOK OF CHEMISTRY

YE OLD ALCHEMIST AT WORK.

THE
BOYS' BOOK OF CHEMISTRY

A SIMPLE EXPLANATION OF UP-TO-DATE CHEMISTRY

TOGETHER WITH MANY EASILY MADE EXPERIMENTS

BY

CHARLES RAMSAY CLARKE

WITH MANY PICTURES BY THE AUTHOR

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Please Read This Preface

How would you like to take two chemicals, mix them together and get a substance entirely unlike either of those you used? or better, take a compound, let the light fall on it, or heat it a little, or pass a current of electricity through it, and change it back into its orginal elements?

If you would like to do either of these things you will certainly like chemistry for this is exactly what this branch of science consists of. And after you have made a few experiments you will have to admit that there is nothing in the whole realm of matter that is more wonderful, more spectacular and, hence, more interesting than chemistry.

Now there are three ways for you to take up chemistry and these are, *first*, to make a lot of experiments just for the fun of the thing, *second* to make some experiments and learn as much as you can about what takes place as you go along, and *third* to carefully make a series of experiments and with the help of this book study the actions and reactions of the chemicals on each other so that you will know exactly how and why they set up the effects and produce the results they do.

The first is the easy, pleasant way, the second is a little harder, though it will not tax your brain overmuch, but the third is the way to really understand chemistry. I have written this book so that you can follow your own bent and choose whichever way may please you best, but my advice is for you to tackle the last mentioned way for you can then make chemistry serve you well—but this will depend somewhat on your age.

Should you simply want to make acids, bases and salts, or decompose solutions and the like, you do not need to pay any attention to the letters which stand for the chemicals, or *symbols* as they are called, and the *formulas* they form which are given. All you have to do is to mix up the chemicals and go through with the operations in the manner explained and you will get the results you want.

But you will find that if while you are making the experiments you will read carefully about the actions that take place between the chemicals, learn the symbols and study the formulas, by the time you get to the end of the book you will have a pretty fair understanding of the subject.

However to get right down to real chemistry you must do all of the above things and then *study*—just mere reading will not do—the chapter on *Practical Chemistry Simply Explained*, and when you know it you and chemistry will be life-long friends. Further, when you are absorbing the laws of chemistry in the chapter just named read the *Appendices* at the end of the book for these also throw light on the names of chemicals which will make things clearer.

I have not aimed to give a complete list of all the chemicals used in the experiments described, for they would cost you too much; but what you should do is to go over the book, check off the experiments you intend to do, and buy the chemicals for these and no more.

Now I believe I here have told you all I can to make you like and understand chemistry so read what I have written first, then fit up your laboratory, get your chemicals and apparatus, and finally make your experiments. You will find that after you have gone this far, which is about halfway, chemistry will come the other half way to meet you and you will thenceforth sail on smooth waters.

CHARLES RAMSEY CLARKE.

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THE BOYS' BOOK OF CHEMISTRY

CHAPTER I

WHEN CHEMISTRY WAS YOUNG

I WONDER if you have ever read of the alchemists of old and of their wondrous art! If not you have missed a story that is far more gripping than the Arabian Nights, or any other tale the Orientals ever told.

Now ever since man became intelligent enough to know that there were such things as earth, air, fire and water he thought over and studied about these marvelous substances long and hard.

It just so happened, though, that the primate we call *man* did not always think rightly—instead he mostly thought wrong—and from his experiments with these *elements*, as he early called them, together with his usually incorrect conclusions he invented what came to be known as the *art of alchemy*. But from out of these experiments and the very few correct ideas that he did gain there sprang the great, modern science that we call *chemistry*.

How Alchemy Came to Be.—If we are to believe the old writers Adam was the first alchemist, and perhaps he was, for, according to the old Testament, he tried to change the forbidden apple into knowledge by eating it. Be that as it may, chemistry did not have its real beginnings until the latter half of the 18th century, that is about the time that our Liberty Bell was ringing out the independence of the United States.

Certain it is that had it not been for the huge blunders and strange conclusions which the alchemists made and deduced from their experiments the science of chemistry would, doubtless, long before then have sprung up and flourished.

But in all the years from the beginning of time to the 19th century the power to think rightly was limited to a very few while most of the people could not reason as well as can a child eight or nine years old of to-day.

Since there were so many discoveries to be made it was quite natural that once in a long while one of the few really bright minds would hit upon something that was of real value to science, while at the same time a hundred foolish ideas were put forth.

Nearly always the former would be rejected by the many because it seemed so commonplace while they gladly accepted the latter in virtue of the fact that they promised such things as changing the baser metals into gold, everlasting life, nearly, etc.

As if this untoward condition of affairs was not bad enough many of the old-time workers along chemical lines who could have discovered and taught real truths taught follies instead since these sharp practices helped them to obtain money in one crooked way or another. The men who thus worked on the ignorance of the people were the patent medicine men of their time and it is to them that, in a large measure, the delay of modern chemistry was due.

What Alchemy Really Was.—Now the alchemy of old consisted of three things and with these the alchemists worked to make something out of little, or nothing, and like impossibilities. Named these things were (1) the *elements*; (2), the *principles* and (3) the *essence*.

In these days of modern chemistry we have over eighty *elements*—that is substances that cannot be decomposed into other substances by any known means—which are in common use and more are being discovered all the time. The alchemists had but four " elements " and even these did not correspond to our definition of an element as you will see presently.

Their elements were *earth, air, fire and water* and when these so-called *elements* acted on one another the three *principles* were formed, or at lesat this is what they so believed. Thus fire acting on air made *sulphur,* air acting on water made *mercury,* while water acting on earth made *salt.*

Hence sulphur, mercury and salt were called the three *principles,* and it was with them and the *elements,* which they thought formed them, that the alchemists worked to get what was known as the *essence,* that is the most important of all things in alchemy.

The Elixir of Life.—They believed that all the substances on earth were composed of but one thing and this was a sort of chemical that they called the *essence.* So vague was the idea, though, of what this wonderful, intangible substance was that nearly every old alchemist had his pet name for it and so the essence was called by such high sounding titles as *The One Thing, The Stone of Wisdom, The Philosopher's Stone, The Heavenly Balm, The Divine Water, The Phoenix, The Old Dragon, The Lion, The Carbuncle of the Sun,* and last and best liked of all, *The Elixir of Life.*

It was believed that if this magic essence could be discovered the alchemist could do with it that which men wanted most and these were (1) to be as rich as Solomon, and (2) to live to be as old as Methuselah. Hence the great reverence in which it was held and the large number of workers who toiled earnestly over their furnaces day in and night out, heating this and distilling that and trying their utmost to *find* the *soul of all things.* But instead of finding their heart's desire they all ended in the same ignoble way and that was with broken health and wrecked fortunes.

The Alchemist's Idea of Alchemy.—The notion on which the old alchemy was based is something like this: when a seed is planted in the ground it decays to all intents and then takes root and grows when a plant springs up.

The alchemists supposed that the metals acted in a like manner, that is if the seed and the method of making it grow could be discovered any metal could be made to grow just as a seed does. Now since everything in *nature* tends to grow toward perfection, the growing metal must do likewise.

Gold, as they thought, was the most perfect metal, hence all other metals must grow into gold if they only waited long enough; but since this would take centuries the alchemists could not wait that long. Their purpose then was not to make gold from nothing but to help lead, iron or mercury grow faster, so that it would more quickly turn into gold. This hurrying up of a supposed natural process was known as *the transmutation of the baser metals into gold.*

The Transmutation of Baser Metals into Gold.—Nearly every old alchemist had a different recipe for making gold but of course none of them ever succeeded in doing so. One of the ways by which they thought it possible to make gold was as follows:

Lead was burned, or *calcined,* as it was called, in a crucible and some of the *essence* dropped into it. In virtue of the marvelous properties that the *essence* possessed it would be found that the lead had turned into gold which was much finer than any found in nature.

The idea that the baser metals were constantly growing more nearly perfect to form silver and gold seemed to be quite certain for when a large piece of lead was calcined in a crucible a mass of cinders was left and in the bottom of the crucible a small button of pure silver was found.

What the old alchemists did not know was that all lead contains a small per cent of silver just as it comes from the mines and that this accounts for the button of silver which was found in the crucible. But even had they known this they would probably have concluded it was due to the supposed fact that the lead was slowly growing into the more perfect silver.

How Alchemy Got Its Name.—I have just told you how the alchemists got a little silver by calcining a lot of lead, which simply means burning it. Now the Arabs have the word *al* which means *the* and the *Greeks* have the word *chemia* which means *calcining*. The old wise men combined the two words and formed the word *alchemy*, which freely translated means *burning*, or, as they would say it, the *art of calcining*.

Some Alchemists of Old.—Nearly all the old alchemists wrote books and each freely described his own favorite method for making gold from the baser metals. The reason they were not afraid of anyone stealing their business was because inasmuch as they themselves could not succeed they knew there was small chance of any outsider doing so.

Among the ancient alchemists who wrote books was *Philalethes* who said much and told nothing in a volume entitled *A Brief Guide to the Celestial Ruby*. Then came *Morton* with his collection of rules which he dubbed *An Ordinal of Alchemy* and finally, *Bonus*, who gave his work a title that sounds very like one of the six best sellers, to wit, *The New Pearl of Great Desire*.

The two outstanding figures in alchemy though, are *Paracelsus* and *Count Cagliostro* (pronounced *ca-lyos'-tro*). Paracelsus was a Swiss doctor whose real name was *von*

Hohenheim. He was born about the time that Columbus discovered America.

He believed, or pretended to believe, that magnetism was a cure for all the ills the human flesh is heir to, and he sold a tincture which he said he extracted from a magnet, but this of course was untrue. But Paracelsus did much good work in the cause of chemistry as you will presently see.

FIG. 2.—A MODERN CHEMICAL LABORATORY.

There were many alchemists who worked faithfully, for they believed in themselves and their schemes but there were numerous others whose sole purpose was to prey upon the ignorant and so get money and enrich themselves.

Such a one was Joseph Balsamo, born in 1743 at Palermo, Italy. He later became known as Count Alessandro Cagliostro. He was a trickster if ever there was one and had long been living on his wits when the idea of taking up alchemy occurred to him.

By clever manipulation he managed to get introductions to several members of the London nobility and who, it so happened, were badly in debt. By means of the most brazen substitution he demonstrated his supposed ability to change lead into gold, collected a large amount of money from his victims to enable him, as he explained to them, to make enough gold to pay off their debts.

This fine story of his sounded mighty good to their hard-up ears but when they had raised and given him the money he demanded he quietly took *French leave*, which means that he suddenly disappeared, and left them *broke* harder than before.

How Chemistry Came from Alchemy.—It was charlatans such as Balsamo who all but threw the gentle art of alchemy into disrepute. However, the larger number of alchemists were not unscrupulous adventurers and triflers but earnest workers who would fain find the essence.

Paracelsus, of whom I told you, was the first to make *hydrogen* and he also discovered *laudanum* and other medicines; by so doing he redeemed himself and his art and this was the real beginning of chemistry. Boyle, who was born in 1626, studied the air and its effects on calcinations, that is burning, or *combustion* as it is called.

Jean Ray, a Frenchman, was the first to show that air was a mechanical mixture of two gases and not a chemical compound. Mayow, who followed, discovered that the air contained *oxygen* and Lavoisier (pronounced *la-vwa-zye*) who discovered the *nitrogen* in the air, were all early chemists who built up the science of chemistry as we know it to-day.

It is due to these wise men who lived in the long ago that the knowledge of chemicals and chemical processes not only makes possible the high state of civilization we now have but also the terrible war which began in the summer of 1914. Had they, like many others of their profession, hidden the truths they learned behind a veil of trickery for the money they could have made out of it, we should still be living the same condition of semi-savagery and ignorance that they then lived in.

And now with this foreword of how the wonderful science of chemistry came to be and its development you are ready to learn a few of the simple but mighty truths of chemistry as we know and use them in our every day lives.

CHAPTER II

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CHAPTER II

THE APPARATUS AND CHEMICALS YOU NEED

LIKE every other branch of science chemistry is based upon certain facts which were discovered by *experimentalists* either through accident or by cleverly working them out.

Now the easiest way for you to learn chemistry is to repeat the *fundamental*, or groundwork, experiments upon which the science is built for then you can see at first hand exactly how chemicals act on one another and what effects these *reactions*, as they are called, will set up.

By experimenting you will get a much better insight into the *how* and the *why* of things chemical than you could possibly obtain by reading, or even studying, a book on the subject. Moreover as you experiment you will run up against a lot of little difficulties which will make you *think* and this is, or at least it ought to be, the real purpose of experimenting.

In this book I have written down a large number of experiments that are not only instructive but many of them are brilliant and fascinating. The apparatus and chemicals you need to perform all of the experiments, nearly, cost very little, the pleasure of doing them is two-fold and the benefits which will accrue to you if you make them will be many and last as long as you live.

Fitting up Your Lab.—Should you want to perform a few experiments for the mere fun of it you won't need a laboratory, but if you intend to make all the experiments which follow with the idea of actually learning chemistry you can't very well get along without one.

Your first move, then, is to get a small room where you can work in quiet and peace. It should be light and airy and the more light and the more air it has the better *lab* it will be. To make it a good place to play in, or work in, as you may care to call it, it should have two other things and these are (1) running water and (2) illuminating gas.

If these fluids are not at hand you will have to carry water to and from your lab every time you want to make an experiment, and use an alcohol lamp whenever you want to heat anything, and both of these are an abomination to a fellow who is trying to do real work.

RIGGING UP A SINK.—The sink should be of good size and of course a porcelain one is the ideal kind. If a high-toned sink like this is out of the question you can buy an enameled iron sink which is just as good, or better, for about $1.50; or you can make one that is plenty good enough by following these directions.

Make a box of 1-inch thick pine boards 10 inches high, 12 inches wide and 3 feet long and bore a 1-inch hole in the bottom for the outlet pipe. Line the box with thin sheet lead or copper and solder all the seams to make them watertight.

Punch half-a-dozen holes through the metal lining over the outlet hole with a sharp-pointed nail; screw the sink solidly to the wall, fix a pipe into the hole and lead it outside of the building where the waste water can run away.

By making a box for a tank and fixing it above the sink you can have a supply of water. To make the sink and a tank is some trouble, to be sure, but you will find it well worth your while in the end.

MAKING A LAB BENCH.—The next thing in order is a bench, or *desk* as it is called in colleges, where you would suppose they would know better.

Make it of 1½ or 2-inch thick boards 2½ feet wide and

5 feet long and have it of such a height that the top of it will be about 2 feet from your eyes when you are standing up. This will give all the room needed for the stands and at the same time you can look down and see into the apparatus you are working with.

One end of the bench should butt up against an end of the sink and be flush with the top of it, as shown in Fig. 3 which also shows how it is put up.

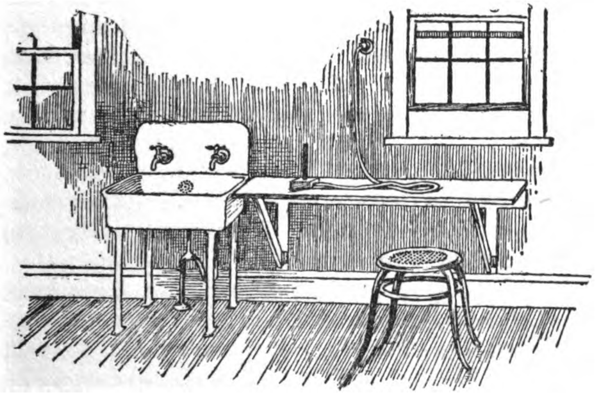

Fig. 3.—A Good Laboratory Arrangement.

If you can afford it you should cover the top of the bench with thin sheet copper, or tin plate will do in a pinch, for this will make it easy to clean up. And by all means clean up when you are through making a series of experiments for working in a messy lab takes away a great deal of the pleasure of it.

The Apparatus and Chemicals.—And now that your playshop is all ready you must get the materials to play chemist with.

There are two kinds of material used in chemistry and these are (1) *the apparatus*, and (2) *the chemicals*. The apparatus consists of the different utensils, devices and tools which you will use during the course of your experiments. The chemicals are the various substances which you will use in combination with the apparatus to make your experiments with.

THE APPARATUS YOU CAN MAKE.—Though it is much more simple to buy your laboratory equipment all ready to use of a dealer, if your purse is thin you can make much of the apparatus at little or no expense.

Besides it is a mighty good scheme to make whatever you can because work of this kind helps to develop your bump of ingenuity and this makes for further resourcefulness on your part. Of course, there are lots of things that you can't make good enough to do the work with, or even make at all, and these you will have to buy with your hard-earned savings.

What a Bunsen Burner Is.—As you know an ordinary gas flame gives off light and if you hold a plate over it the latter will soon be covered with a layer of *soot.*

Now this substance which we call soot is, in the language of chemistry, *carbon,* and it is the same stuff as the charred part of a stick of wood, though it is slightly different in form.

For heating purposes you must have a flame which is not only hot but which gives off no soot, or carbon, and this is just what a Bunsen burner[1] does. A Bunsen burner is a very simple device and consists of a metal tube or pipe about $\frac{3}{8}$ of an inch in diameter; a rubber tube is connected with a source of illuminating gas and to the lower end of the pipe forming the burner; and, finally, a couple of holes are drilled through the pipe near the lower end so that air can flow into and mix with the gas before it is burned.

This extra air supplies enough *oxygen* to burn up the carbon in the gas and this makes the flame hotter and at

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[1] Invented by Bunsen, a German scientist who lived from 1791 to 1860.

the same time it makes it *non-luminous*, that is it does not give any light.

How to Make a Bunsen Burner.—To make a Bunsen burner get a piece of iron pipe ½ an inch in diameter outside measurement.

About 1½ inches from one end drill a ¼-inch hole clear through the pipe and have threads cut on the end nearest the hole. Next get a block of wood 1 inch thick and 3 inches square and screw a strip of wood 1 inch square to each side as shown in Fig. 4.

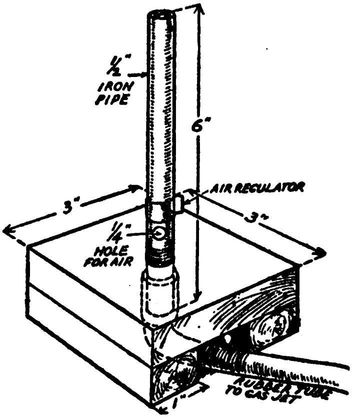

Fig. 4.—How a Bunsen Burner is Made.

Bore a ½ inch hole in the center of the block and screw the threaded end of the pipe into this hole until it sticks through the side with the strips on it ½ an inch. This done, cut out a strip of tin ½ an inch wide and 2 inches long; form a ring of it which will slide on the pipe snugly, and slip it down over the holes, when your burner is complete.

Set it on your bench and fit one end of a rubber tube on the lower and threaded end of the pipe and connect the other end with a gas-jet from which you have removed the tip.

Experiments with a Bunsen Burner.—Which Show the Nature of Flame. Turn on the gas and hold a match just to one side of the top of the pipe when it will light. This is the right way to light a gas jet, for if you hold the match directly in the path of the gas the force of the latter will usually blow it out.

You will please take notice that when the tin ring is pushed down over the holes in the pipe that the flame is *luminous*, that is it gives off light just like an ordinary gas flame when of course it will deposit carbon on anything that is held over it.

You will also observe that when the ring is pushed up and the holes are opened the flame becomes *dark*, or non-luminous, that it no longer produces carbon, and that it is much hotter. This is the kind of a flame you should always use when heating substances. The size of the flame can be regulated by the valve of the gas jet.

The Luminous Flame of the Burner.—Before you go any farther try this little experiment with the burner.

Push the ring down over the holes and get a luminous flame. Study the flame carefully and you will see that there are four separate and distinct parts to it; these parts of the flame, or *cones* as they are called, are,

(1) A dark cone right above the mouth of the pipe; hold the head of a match in this cone and you will find that it will neither light nor burn. This is because this cone of gas is not burning. As a matter of precise statement this dark cone is not a part of the flame at all.

(2) Around this dark cone is a deep blue cone which is pretty hot and you can prove it by holding a match in it when it will instantly *ignite*. It is in this blue cone that the gas first begins to burn.

(3) Above the blue cone is a bright yellow cone and this is known as the *luminous cone*. This cone is quite hot and its color is caused by the particles of carbon which are heated until they *emit*, that is give off, light.

(4) Finally above the yellow cone is yet another enveloping it and which is nearly colorless and this cone is very hot. The cones of luminous flame are shown at *A* in Fig. 5.

The Non-luminous Flame of the Burner.—Having made the above observations push the ring up so that air will mix with the gas and make a non-luminous flame. You will find that the luminous cone has disappeared but that the other cones are still there.

The hottest part of the flame now is the tip and the next hottest is the middle cone while as before the bottom cone is not burning at all. To prove this put one end of a small glass tube into this cone, hold a match near the other end of the tube and the gas will ignite at that end.

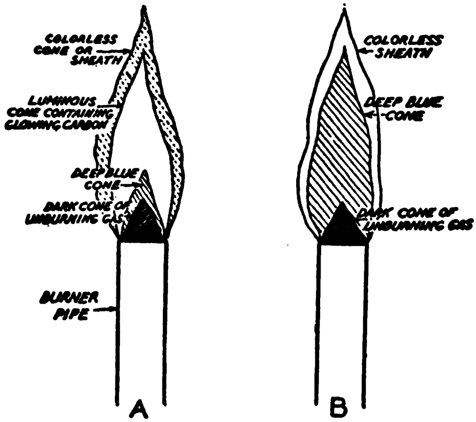

FIG. 5.—A. COMMON GAS BURNER FLAME. B. BUNSEN BURNER FLAME.

The cones of a non-luminous flame are shown at *B* in Fig. 5.

How to Make a Test-tube Stand.—The purpose of a *test-tube stand* is, as you can easily guess from its name, to hold your tubes when you get them.

You can easily make a test-tube stand, and lots of other like things, of *annealed* iron wire, which means that the wire has been heated red-hot and allowed to cool slowly when it loses its *elasticity*, or springiness; it can then be bent to shape without trouble.

Use $\frac{3}{32}$-inch thick iron for the test-tube stand and to bend it neatly you will need a pair of 5-inch flatnose side-cutting pliers—these are absolutely necessary—and one pair of 5-inch round-nose pliers—which you must have if you are to do a good job.

Cut off a piece of wire about 6 feet long and with your fingers and pliers mixed with a little brains bend it to the size and shape shown in Fig. 6. The rings formed by the wire should be about ¾ inch in diameter at the top and about ⅜ inch at the bottom. This stand will hold your test tubes in an upright position side by side and you can make your tests and take notes without bother.

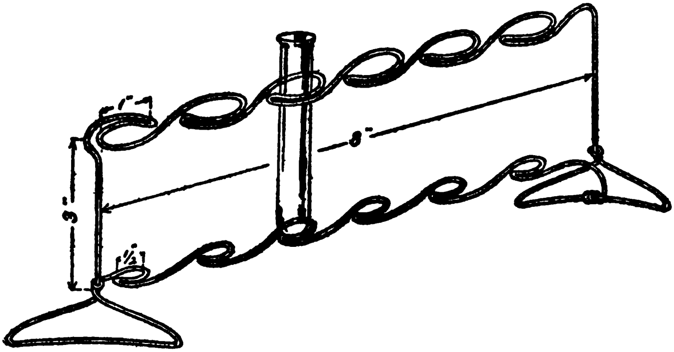

FIG. 6.—HOW TO MAKE A WIRE TEST-TUBE STAND.

Making a Ring Stand.—The next thing is to make a ring stand of $\frac{3}{16}$-inch thick rod. This is a very easy piece of apparatus to make, as you will see by looking at Fig. 7, which has the sizes marked on it.

The lower ring which forms the base is about 8 inches in diameter, and this with the support, which is about 12 inches high, is formed of a single length of rod.

Make two separate rings of ⅛ inch thick wire and have one of them 2 inches in diameter and the other 3½ inches in diameter. Look at Fig. 7 and you will see how these rings are made. You can slip the rings up and down on the standard to whatever height you wish and set a flask in one of them with the Bunsen burner underneath and heat some secret compound in it to a romantic glow, as did the alchemists of old, and gloat over the successful issue of the experiment, which they did not.

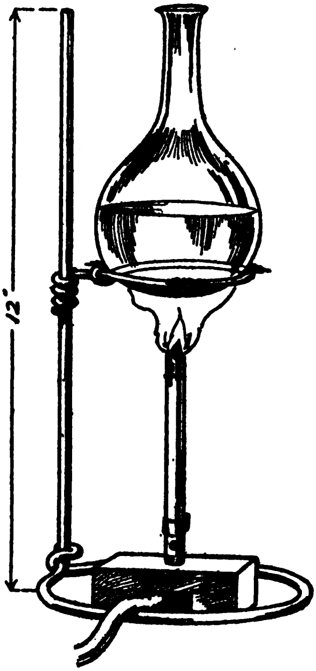

Fig. 7.—The Ring Stand and How it is Used.

Buy a piece of heavy wire gauze about 4 inches square and lay this on the ring and then put your flask or dish on top of it, as this will keep the flame from touching the glass, for while heat will pass through the gauze a flame will not.

A Simple Deflagrating Spoon.—A *deflagrating spoon* is used to lower a substance into a long-necked bottle when you want to see how it burns in some gas or other substance.

To make a deflagrating spoon take a piece of $\frac{3}{16}$-inch wire about 18 inches long as shown at *A* in Fig. 8; flatten one end a little and drill a $\frac{3}{32}$-inch hole through it; this is for the handle.

The spoon proper is made of a piece of sheet iron 1 inch square; cut the corners in $\frac{3}{8}$ inch as shown at *B* and then bend up the edges to make a miniature pan of it. Fasten this pan or spoon to the handle with a screw having a nut on it and you will have added another piece of apparatus to your list.

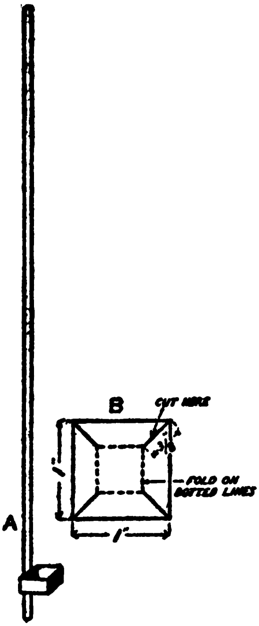

Fig. 8.—The Deflagrating Spoon.

An Easily Made Test-tube Holder.—A test-tube holder is a device that you must have; it is a tweezer-like affair for picking up hot test tubes without burning your fingers.

Take a piece of spring brass wire $\frac{3}{32}$ inch in diameter and bend it over in the middle; twist it around a pipe or rod ½ inch in diameter, forming a couple of turns; this done, bend the rest of the wire to the shape shown in Fig. 9.

Now when you squeeze the holder near the turns of wire at the back the jaws, or free ends, will open, and conversely, when you ease up on the pressure the jaws close tight enough to hold a hot test tube or other object firmly.

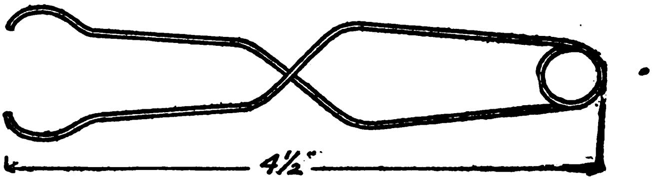

Fig. 9.—A Test-tube Holder.

The Apparatus You Must Buy.—The rest of the apparatus you need consists of glassware chiefly, and while it is better to buy all of it you might be able to use utensils to be found in the kitchen for some of them.

The apparatus needed is: (1) three *beakers,* one of which is 2 inches in diameter and each of the other two having a diameter of 3 inches. Instead of beakers ordinary glasses can be used where heat is not needed.

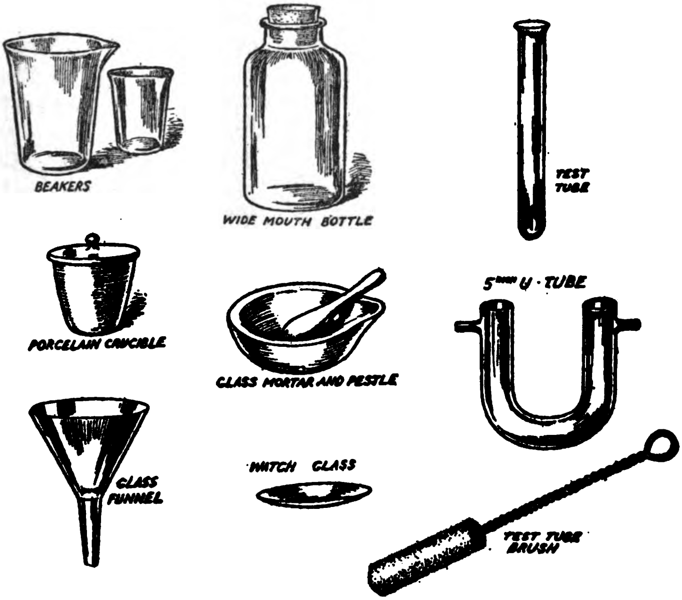

Fig. 10.—The Apparatus You Will Need.

Beakers are used for holding liquids and mixing solutions so that they can be examined and experimented with.

(2) Three wide-mouthed bottles with corks. Fruit jars will serve the purpose if they are of good clear glass; they are used for collecting and holding gases; (3) one dozen 5-inch test tubes for examining and testing small quantities of liquids and solids, and get one dozen corks to fit the tubes.

(4) One porcelain crucible and cover for melting solids; (5) two porcelain evaporating dishes for evaporating liquids over a flame; (6) one glass mortar and pestle for grinding and crushing solids; (7) one 5-inch U-tube for drying and purifying gases; (8) one glass funnel to aid in pouring and for filtering liquids.

(9) Two 2-inch watch glasses for evaporating solutions without heat; (10) six 2-foot lengths of glass tubing, standard size, for connecting different pieces of apparatus together; (11) two packages of *litmus paper*, one red and one blue, for detecting acids and salts.

(12) One piece of platinum wire 4 inches long for aiding reactions between acids and metals, as you will presently see; (13) one package of filter paper for filtering liquids and solutions; (14) one 5-inch *three-square* (triangular) file for cutting glass tubing; (15) a couple of feet of rubber tubing for making connections between apparatus; (16) one test-tube brush for cleaning test tubes; (17) two towels and a cake of soap for cleaning up, and, lastly, (18) a sponge for cleaning your bench. The chief things which I have just described above are shown in Fig. 10.

The Metric System of Weights and Measures.—You may not know it, but in chemistry the *metric system* is used instead of the English system to measure out substances.

The metric system is much more simple than our English system of weights and measures, and so you don't need to be afraid to tackle it. In all European countries the metric system is used altogether and it ought to and doubtless it will be used here in every-day life before long.

Now in chemistry the metric system of weights and measures is the only one, and if you are going to be a real chemist you must learn to know it just as well as you do the English system. In the latter system we use the *inch* and fractions of the inch, the *foot* and the *yard* when we measure length.

In the metric system the measures of length are the *millimeter*, the *centimeter*, the *decimeter* and the *meter*. The relation of these measures to the English system is shown in Fig. 11, and the relation of the latter to each other is given in the following table:

TABLE OF METRIC MEASURES OF LENGTH

10 millimeters = 1 centimeter;
10 centimeters = 1 decimeter·
10 decimeters = 1 *meter*;
or 100 centimeters = 1 meter
or 1000 millimeters = 1 meter.

Note.—1 millimeter is just about $\frac{3}{64}$ of an inch long; 1 centimeter is about $\frac{7}{16}$ of an inch long and 1 meter is nearly 40 inches or to be exact 39.39 inches long.

It takes only half an eye to see that the metric system which runs in tens is much more simple than the English system, which is based on neither rime nor reason.

In the metric system *volume*, that is, liquids and other substances in *bulk*, is measured in *cubic millimeters*, *cubic centimeters*, *cubic decimeters*, and *cubic meters*.

TABLE OF METRIC MEASURE OF VOLUME

10 cubic millimeters = 1 cubic centimeter;
10 cubic centimeters = 1 cubic decimeter;
10 cubic decimeters = 1 cubic meter;
or 10 cubic decimeters = 1 *liter* (pronounced *lee-ter*);
or 1000 cubic centimeters = 1 liter.

Knowing now what the metric system of measures is you will be the better able to understand the nature of the following apparatus when you come to buy it.

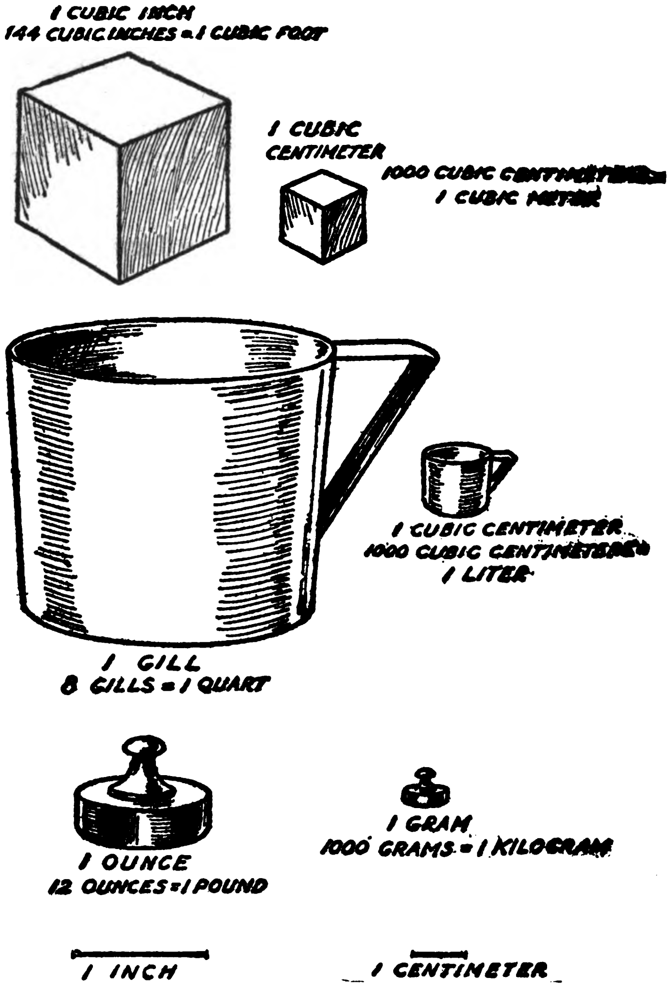

FIG. 11.—ENGLISH AND METRIC MEASURES COMPARED.

[*These drawings are ¾ of full size.*]

The Graduated Cylinder and How to Use it.—Buy a 25-cubic centimeter—or c.c., as it is called for short—graduated cylinder. This is merely a tall glass which has a marked scale on it reading up from 1 c.c. to 25 c.c., on the outside. It is used to measure any quantity of a liquid up to 25 c.c., and this is done by simply pouring it in until it reaches the mark indicating the amount you want.

You will also need two *Florence flasks*[1] one of which will hold 100 c.c., the other 250 c.c., and both should be fitted with rubber stoppers. These flasks are round bottles of thin glass and are used for heating liquids and solutions.

Finally get a wide-mouthed bottle holding 1 liter and a piece of ground glass 8 centimeters square. The purpose of the bottle is to collect samples of gases for weighing and the ground glass is used for a cover to keep in the gas that is contained in the bottle.

About Weights and Scales.—Next is the metric system of weights. Now the *weight* of 1 cubic centimeter of water at 4 degrees *centigrade*[2] is called 1 *gram* and this is the metric unit of weight.

TABLE OF METRIC MEASURE OF WEIGHTS

1000 milligrams = 1 gram;
100 centigrams = 1 gram;
10 decigrams = 1 gram;
1000 grams = 1 kilogram and
1 liter of water weighs 1 kilogram.

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[1] So called because this kind of a flask originally came from Florence, Italy, and contained olive oil.

[2] *Centigrade* means graduated to a scale of 100. A *centigrade thermomenter* is used in scientific experiments and on its scale the freezing-point of water is 0 and the boiling-point of water is 100. See Appendix A.

To experiment the right way with chemicals you should have a set of weights which will weigh down to at least a *centigram*, or better yet to a *milligram*. See Fig. 12.

You must also have a *scale* or *balance*, as it is sometimes called, which the makers guarantee will weigh accurately to a centigram and such a scale can be bought for about $4.00. You can of course use a cheaper scale if you are making the experiments for the fun of the thing rather than for the knowledge you will gain.

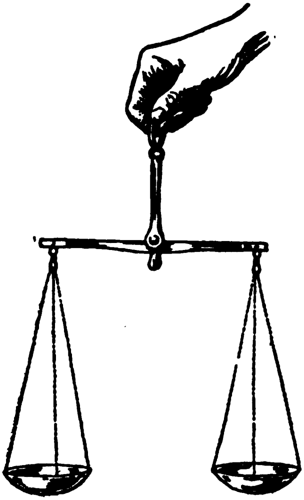

FIG. 12A.—A SIMPLE HAND SCALE.

Your Laboratory Note Book.—You will find that after doing an experiment the details will quickly slip from your mind unless you jot them down as you go along.

Now since the experiments you make will be of small value unless you are able to recall them you should by all means write them up in a note-book and the handiest of these is of the *loose-leaf* kind.

The Chemicals you Need.—Last of all are the chemicals and the following list will be sufficient and complete enough for you to perform all the standard experiments in this book:

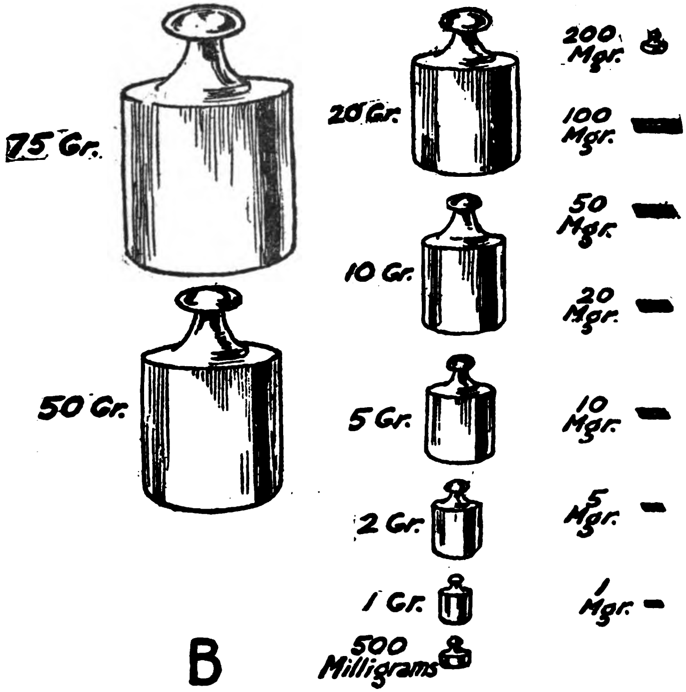

Fig. 12B.—Metric Weights from 75 Grams to .001 Gram or 1 Milligram.
[*Approximately.*]

Amount.	*Name of Chemical.*
16 oz. bottle	hydrochloric acid, conc. sol.
16 oz. bottle	sulphuric acid, conc. sol.
8 oz. bottle	nitric acid, conc. sol.

Amount.	*Name of Chemical.*
16 oz.	sodium hydroxide solution, conc.
4 oz.	ammonium chloride.
16 oz.	sodium chloride.
4 oz.	carbon disulphide, conc. sol.
8 oz.	animal charcoal.
8 oz.	copper foil.
4 oz.	cupric sulphate crystals.
1 oz.	cupric oxide.
4 oz.	fluorspar.
1 oz.	iodine, sublimated crystals.
8 oz.	starch.
4 oz.	iron filings (fine).
8 oz.	ferrous sulphide.
4 oz.	ferrous sulphate.
2 oz.	lead nitrate, conc. sol.
2 oz.	silver nitrate, conc. sol.
16 oz.	manganese dioxide, powdered.
1 oz.	red phosphorus.
2 oz.	potassium bromide.
8 oz.	potassium chlorate.
2 oz.	potassium bichromate.
1 oz.	potassium iodide.
2 oz.	potassium permanganate.
4 oz.	potassium nitrate.
2 oz.	Glauber's salts.
2 oz.	anhydrous sodium sulphate.
4 oz.	sodium nitrate, conc. sol.
4 oz.	sodium sulphate.
8 oz.	roll of sulphur.
16 oz.	granulated zinc.
2 oz.	zinc sulphate.
4 oz.	mercury.

All the chemicals listed above will come in solid form except those marked *conc. sol.*, which means that it is a *concentrated solution.*

It is the better way to have all of the solid chemicals put up in bottles, or *phials* as they are sometimes called. These bottles will cost from 2 to 8 cents extra per chemical and they are more than worth it for they will effect a great saving, keep the substance in good condition and have printed labels on them so that you can see just what you have at a glance.

I have made no attempt to quote the cost of these chemicals because the prices are constantly changing and usually tend toward a higher mark.

When you are all ready to buy them send in your list to the L. E. Knott Apparatus Company, Boston, Mass., or to Eimer & Amend, 205–211 Third Avenue, New York City, and either firm will quote you the right prices.

CHAPTER III

EXPERIMENTING WITH WATER

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EXPERIMENTING WITH WATER

WATER, as every red-blooded American boy knows, is a chemical compound that is to be used very sparingly on the hands and face but is great to get into on a hot summer's day down in the ole swimmin' hol'!

Next to the air we breathe water is the most important of all the things which enable us to live here on this hot earth of ours. Since this is so it is a mighty good thing that three-fourths of the earth's surface is covered with this useful fluid.

Though everything you eat or drink contains water, and though you yourself are composed largely of this liquid substance I dare say you have never stopped for a minute to think about what water really is and how it is made.

How Water Looks.—Draw a beaker of water from the tap, hold it between your eyes and the light and you will see that it is nearly colorless. Now think back to the way the old ocean looked the last time you saw it and if your memory serves you well you will picture in your mind's eye that it was blue.

The question that comes up is, why is the water in the beaker colorless while that of the ocean is blue? The answer is easy; water really has no color but when you look at a large quantity such as a lake or the sea you see in it the reflection of the sky and clouds above it, or the effect of materials or chemicals which it contains.

Water is a Compound.—But water is much more than just a colored or colorless liquid as you chance to see it. It is a *compound*, that is it is a substance formed chiefly of two substances, or *elements*, as they are called, in combination with each other.

About Mixtures and Compounds.—EXPERIMENT SHOWING WHAT A MIXTURE IS.—In the good old days before pure food laws were invented and the unscrupulous grocer mixed a little sand with a lot of *grade C* sugar he got a simple *mechanical mixture*, and so will you if you try it, that is the particles of sand and sugar are not changed in any way.

FIG. 13.—GRINDING MERCURY AND SULPHUR IN A MORTAR.

A mixture, then, is a *physical process* in which the substances that are mixed remain the same as before they were jumbled together.

EXPERIMENT SHOWING WHAT A COMPOUND IS.—Pour a few drops of *mercury*, which is a metallic *element*, into your mortar and then add a little *powdered roll sulphur*, which is also an element.

Grind these two elements together in your mortar with the pestle, as shown in Fig. 13, and in a few moments you will have instead of the bright mercury and the yellow sulphur a putty-like black mass, the chemical name of which is *mercuric sulphide* (HgS) and which looks like neither the one nor the other of its *component* elements. The two elements have united to form a *chemical compound.*

Further while it is easy to separate the different substances which form a simple mixture it is very hard indeed to separate the elements which have chemically combined to form a compound.

It may interest you to know that mercuric sulphide when it is heated forms a red powder, called *red mercuric sulphide,* and this is sometimes used instead of red lead in house paints because it is more lasting.

Water is Formed of Two Gases.—Though water may seem to be a simple substance it is formed of two gases and these are called *oxygen* (O) and *hydrogen* (H). And now before I tell you how to combine these two gases to form water let's find out the important things about them.

Oxygen, the Magnetic Gas.[1]—WHERE IT IS FOUND.—Though oxygen is a common gas it is none the less wonderful and you must get acquainted with it.

It is found everywhere and in nearly everything. Here are a few facts about this gas that may sound a trifle *gaseous* but they are true, every one of them: (1) one-fifth of the air, measured by volume, is oxygen; (2) eight-ninths of water is oxygen, and (3) one-half of the solid crust of the earth is formed of oxygen. So you see it is an all-important gas to we mortals who live on this little planet.

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[1] Oxygen is the only known gas that can be magnetized.

Experiment Showing How to Make It.—The best way to get acquainted with oxygen is to make, examine and experiment with it yourself.

Look over your chemicals and get out the bottle labeled *potassium chlorate* ($KClO_3$). This is a compound formed of white crystals and is largely used in making matches and fireworks.

This chemical is made up of 1 part of *potassium* (K), 1 part of *chlorine* (Cl) and 3 parts of oxygen (O).[1] Put a little potassium chlorate in a test tube and hold it in the flame of your Bunsen burner until it melts and finally boils.

Take a match, light it and after it is blazing blow it out; stick the glowing end into the test tube when the splint will burst into a flame. This experiment shows that there is oxygen in the test tube.

Moreover it is a test for oxygen for there is no other gas[2] that will make a glowing splint light after the flame is once blown out.

The Experiment Simply Explained.—The way the oxygen came to be in the test tube is like this: when you put the potassium chlorate over the flame the heat drove out the oxygen that was in it and this left the potassium and the chlorine behind in the form of a white salt.

This salt is called *potassium chloride* (KCl), so you see that potassium chlor*ide* is simply potassium chlor*ate* minus the oxygen.

The Sign Language of Chemistry.—A little way back I told you that potassium chlorate was made up of 1

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[1] Why potassium chlorate has 3 parts of oxygen is explained in Chapter VIII, on Practical Chemistry Simply Explained.

[2] As a matter of exact statement there is one other gas which supports combustion and this is *nitrous oxide*, or *laughing gas* as it was once called. It is composed of 2 parts of nitrogen and 1 part of oxygen, or N_2O, to write down the formula in *symbols*. For all reaction purposes, however, the glowing splint is used as a test for oxygen.

part of potassium, 1 part of chlorine and 3 parts of oxygen.

Now instead of writing out the full name of a chemical we can make it much clearer if we write it so that we can see at a glance exactly what these elements are or how these elements combine. Of course we can do this by writing out each substance and its number of combining parts:

Thus for potassium chlorate we can write, *potassium, chlorine, oxygen* 3, the 3 showing that there are 3 parts of oxygen in it; since there is only 1 part of potassium and 1 part of chlorine we do not need to take the trouble to write 1 after them for this is understood.

But as it is often hard to write out the full names of the elements in a compound it has been agreed upon by chemists to use letters or signs, called *symbols*, instead. Now the symbol for potassium is K, that for chlorine is Cl and that for oxygen is O.

As you come to the symbols you should remember them for they are used by chemists and are found in all books on chemistry. So instead of writing out the full names of the elements forming potassium chlorate as above stated you need only to write the symbol for it which is $KClO_3$.

You can also show what takes place when a chemical compound is heated or when one chemical acts on another by writing down the symbols in this way:

$$\underset{\text{Potassium chlorate}}{KClO_3} + \text{Heat} = \underset{\text{Potassium chloride}}{KCl} + \underset{\text{Free oxygen}}{3O}$$

Since you had 3 parts of oxygen to start with on the left-hand side of the equality sign[1] (=) you must have

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[1] An equality sign shows that the two things on the opposite sides of it are equal. An arrow pointing in the direction in which the action goes is now used in modern chemistry to take the place of the old equality sign.

3 on the right-hand side of it while the number of parts of potassium and chlorine also remain the same. When symbols are used as above they form what is called an *equation*.

Making a Large Amount of Oxygen.—In the course of your experiments you will need a considerable amount of oxygen.

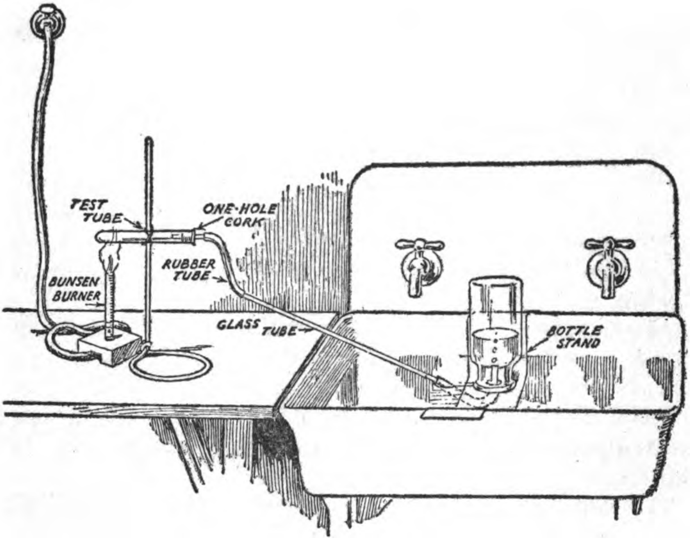

FIG. 14.—HOW TO SET UP THE OXYGEN APPARATUS.

To make oxygen on a fair scale set up your apparatus as shown in Fig. 14. Next weigh out 5 grams of *potassium chlorate* and 3 grams of *manganese dioxide* (MnO_2) and mix and grind them thoroughly in your mortar; put the mixture into a test tube and fix the tube to your ring-stand with a rubber band.[1]

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[1] When fixing a test tube to a ring stand in this fashion to be heated always have it slant downward to the mouth of

Nearly fill the sink with water and set the bottle stand described in the last chapter on it as shown in Fig. 14. Fill a wide-mouthed bottle clear up to the mouth with water and turn it upside down taking care to keep the mouth under the water in the sink so that the outside air can't get into the bottle.

Set the mouth of the bottle down on the bottle stand over one of the holes. This done fix up an arrangement of rubber and glass tubes as is also shown in Fig. 14 and put the end of the tube through the hole in the bottle stand and into the mouth of the bottle.

Heat the test tube quite hot until all the water in the bottle has been driven out or displaced by the oxygen. Then slip one of the ground-glass plates under the mouth, hold it on tight, take the bottle out of the sink as shown in Fig. 15, turn it right side up and put it on your bench; you will then have a bottle full of oxygen.

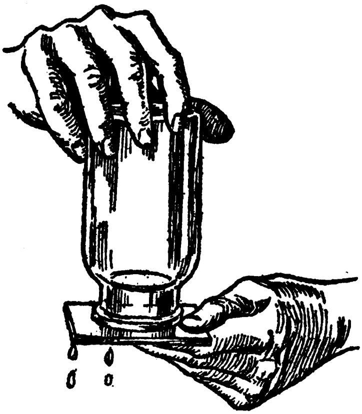

FIG. 15.—REMOVING THE BOTTLE OF OXYGEN FROM WATER.

How This Chemical Action Works.—In making oxygen the manganese dioxide that you put into the test tube did not give off any oxygen but it *helped* the potassium chlorate to give it off faster.

Any substance which has this peculiar property of

the tube; if then there is any water in the chemicals and the heat drives it out it will not run back into the heated part of the tube and crack it but will run down into the mouth which is cool.

aiding a reaction without being changed itself is called a *catalytic agent*,[1] thus:

$$\overset{\text{Potassium chlorate}}{KClO_3} + \overset{\text{Manganese dioxide}}{MnO_2} = \overset{\text{Potassium chloride}}{KCl} + \overset{\text{Manganese dioxide}}{MnO_2} + \overset{\text{Oxygen}}{3O}$$

You will see from this equation that the manganese dioxide remains unchanged and that the oxygen in the potassium chlorate is set free.

FIG. 16.—SULPHUR BURNING IN OXYGEN.

Some Experiments with Oxygen.—AN EXPERIMENT IN DEFLAGRATION.—Hold the bottle of oxygen between your eyes and the light and the bottle will appear perfectly empty; this is because oxygen is colorless; you don't need to test it but it is also odorless.

For this experiment put a piece of sulphur in your deflagrating spoon and light it by holding it in the Bunsen flame; you will observe that it burns with an infernal odor and emits a feeble but none the less diabolical light. This is the *brimstone* of Bible lore and it is no wonder that the ancients had a wholesome respect for it.

Now lift the ground glass from the bottle of oxygen and stick the spoon down into it as shown in Fig. 16.

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[1] A *catalytic agent* is sometimes called a *catalyst* or *contact agent*. *Catalytic* comes from two Greek words which mean *down* and the act of losing, or *losing down*.

Instantly the sulphur burns with a much brighter flame, and you can then smell the odor of it. When you read Chapter V you will learn why the sulphur burns and smells differently in air and in oxygen.

EXPERIMENT WITH A SCINTILLATING WATCH SPRING.—Fill another bottle with oxygen; straighten a piece of the mainspring of a watch and fix a bit of sulphur on one end of it. Light the sulphur and let it down into the gas when it will burn with a brilliancy like unto an electric arc light.

The Chemical Properties of Oxygen.—From the above experiments you will have gathered that oxygen is the gas which makes things burn.

Forsooth, if it were not for oxygen we could not build a fire, or breathe, or anything, for without it the earth would be as cold and as dead as the moon. In the first experiment when the sulphur burned the heat and the flame were caused by the oxygen combining with the sulphur[1] and it did so rapidly, because the sulphur had already been lit.

But things do not need to burn to make oxygen combine with them though the action is much slower. For instance when iron *rusts* it is simply because the oxygen of the air has combined with the iron to form a new substance or compound. The chemical name for this compound is *iron oxide*, or *ferric oxide* (Fe_2O_3) but its common, everyday name is just plain *rust*.

Rusting is really a kind of slow burning, though burning, or *combustion* as it is called, consists of the combining of a substance with oxygen so rapidly that heat and sometimes flame are produced.

Hydrogen the Lightest Substance Known.—WHERE IT IS FOUND.—As I have told you before, oxygen is only one of the gases that make up the compound we know

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[1] See Chapter V for the reason.

as water. The other gas which forms it is called *hydrogen* (H).

There is a lot of hydrogen in all kinds of animal and vegetable compounds and when you drink a glass of water you are merely taking $\frac{1}{9}$ part of hydrogen by weight into your stomach, *via* your mouth, the other $\frac{8}{9}$ being oxygen.

Setting up the Apparatus.—Weight out 10 grams of *granulated zinc* and gently drop it into your 100 c.c.[1] flask. Next push one end of your funnel[2] into one of the holes of a two-hole rubber cork until it sticks out $\frac{1}{2}$ an inch on the other side.

Now cut off a piece of rubber tubing 2 inches long and slip it over the end of the funnel. Fit the cork into the flask and cut the piece of glass tubing[3] just long enough so that when you fit it into the other end of the rubber tube it will reach within $\frac{1}{2}$ an inch of the bottom of the flask.

Make an L bend[4] in a 6-inch piece of glass tube and fit this into the other hole of the cork. Cut another 2-inch piece of rubber tube and push this on the end and, lastly, slip a piece of glass tube 10 inches long into the other end of the rubber tube as shown in Fig. 17.

How to Know an Apparatus is Air-tight.—When you have the apparatus all set up fill the flask half-full of water and blow into the end of the 10-inch tube.

When the water half-fills the funnel, which it will do if you blow hard enough, clap your thumb tightly over the end of the 10-inch tube, or *delivery tube* as it is called,

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[1] Abbreviation for *cubic centimeter*.

[2] Always wet the cork and soap the glass tube before you push a tube through the hole in it as this makes it much easier to work it into place.

[3] Directions for cutting glass tubing will be found at the end of this chapter.

[4] See also end of this chapter for instructions in bending glass tubing.

and see if the water in the funnel remains there or starts to seep out.

If it stays there all well and good but if it begins to leak out tighten the cork in the flask until it stops. When the water stays in you will know that the apparatus is air-tight and woe be unto you if you fail to make it so.

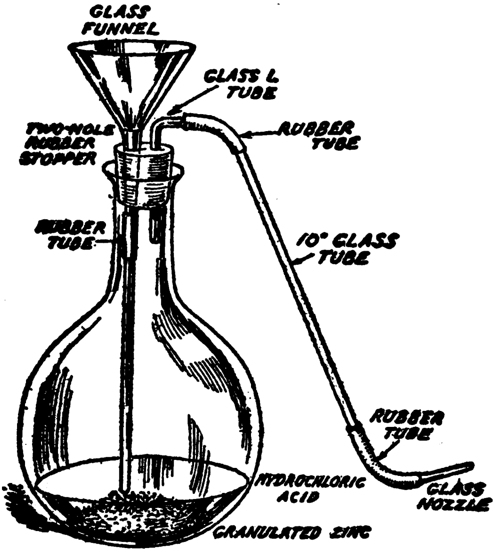

Fig. 17.—The Hydrogen and Water-making Apparatus.

After you have tested the apparatus slip a piece of rubber tubing 2 inches long on the delivery tube and slip the end of a glass nozzle[1] into the other end of the rubber tube, as shown in Fig. 17, when the apparatus is ready to generate hydrogen.

Making Hydrogen Experimentally.—To make enough hydrogen for experimental purposes take 20 c.c. of

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[1] How to make a glass nozzle is told at the end of this chapter.

hydrochloric acid (HCl) in your graduated cylinder and add it to 20 c.c. of water, stirring them well with a glass rod and keeping it off of your hands.

Next pour the water out of your flask and pour about 5 c.c. of the hydrochloric acid solution into the funnel at a time, but wait a few seconds between each time you pour until it is all in and cover the top of your funnel tight with the ground glass when you are not pouring. The instant the acid touches the zinc bubbles of gas are given off and this gas is *hydrogen*.

HOW TO TEST HYDROGEN GAS.—To find out whether hydrogen gas is pure or not let the acid act on the zinc for a couple of minutes and then hold a test-tube mouth downward over the end of the nozzle for a few seconds.

Close the mouth of the test tube with your thumb and carry it, still mouth downward, over to your Bunsen burner which you must move well away from the hydrogen apparatus. Take your thumb away and the gas will light with a sharp noise like the bark of a sky-terrier.

This noise shows that you have not got pure hydrogen in the test tube but a mixture of hydrogen and air which is very explosive. This is the reason why you must be sure to make the apparatus air-tight.

As soon as the hydrogen in the tube has forced out all of the air in it the gas that then comes from the nozzle will not explode when you test it for the hydrogen is pure, and therefore not explosive.

WARNING.—You must take the greatest pains to see that the hydrogen you make is pure or else the apparatus may blow up.

A Wet Flame.—AN EXPERIMENT SHOWING THE CHEMICAL PROPERTIES OF HYDROGEN.—When you are sure that you have pure hydrogen, light the gas coming from the nozzle of the delivery tube; it is a good scheme to wrap a towel around the flask to prevent pieces of glass from flying about should it take a notion to blow up.

You will see that the hydrogen burns with a pale blue and an almost colorless flame. Things will not burn in hydrogen but it will burn itself while oxygen acts the opposite way.

Now as I have told you before, burning is simply the combining of a substance with oxygen and this is just what the hydrogen is doing, that is to say the hydrogen you have made and the oxygen of the air are combining and form a compound. Now the only compound that these two gases can form is water; hence you are *making water*.

To see the water hold a cold, dry beaker in the flame when the bottom of the beaker will collect the water and which, when there is enough, will drip off of it.

How the Experiment Works.—The way hydrochloric acid and zinc form hydrogen is this: The hydrochloric acid is made up of hydrogen and chlorine as you can tell from its name and its symbol HCl.

Now zinc has the power of driving the hydrogen out of the dilute hydrochloric acid and of then combining with the remaining chlorine when it forms *zinc chloride* ($ZnCl_2$). The symbol for zinc is Zn and to make the above action clearer you can write it down in symbols thus:

$$\overset{\text{Hydrochloric acid}}{2HCl} + \overset{\text{Zinc}}{Zn} = \overset{\text{Zinc chloride}}{ZnCl_2} + \overset{\text{Free hydrogen}}{2H}$$

This kind of an equation is also called a *formula* and chemists usually leave out the equality sign and use an arrow[1] in its place like this:

$$\overset{\text{Hydrochloric acid}}{2HCl} + \overset{\text{Zinc (make)}}{Zn \rightarrow} \overset{\text{Zinc chloride}}{ZnCl_2} + \overset{\text{Hydrogen}}{2H}$$

And further since the hydrogen is a gas and escapes from

---

[1] If you take the arrow to mean *make* it will read with good sense.

the compound as soon as it is made a little arrow pointing up is used to show it thus:

$$2HCl + Zn \rightarrow ZnCl_2 + 2H \uparrow$$

Now when you burn hydrogen it combines with the oxygen of the air to form water in this manner:

$$\underset{\text{2H}}{\overset{\text{2 parts of hydrogen}}{}} \quad 2H + \overset{\text{Oxygen of the air}}{O} + Heat \rightarrow \overset{\text{Water}}{H_2O}$$

and hence H_2O is the formula for water. In other words two parts by volume of hydrogen have combined with one part of water. See?

To Prove the Experiment.—You can easily prove that water is formed of hydrogen and oxygen by breaking up, that is *decomposing*, some water and applying the usual tests for hydrogen and oxygen to the gases which are given off. This is done by means of a simple electrochemical process which you will find in Chapter IX under the heading of *The Electro-Chemical Arts.*

Chemicals that Give Off Water.—When a chemical absorbs water or gives it off it is called a *hydrate* and don't forget it either. When some compounds are formed they combine with water and this is often the moisture that is in the air. Thus *cupric sulphate crystals* ($CuSO_45H_2O$), commonly called *blue vitriol*, contain a lot of water.

Making Crystals from Water.—EXPERIMENT WITH WATER OF CRYSTALLIZATION.—Half fill an absolutely clean crucible with the blue crystals and place it on your scales and note with exceeding care its exact weight.[1]

Now set the crucible on the ring stand, with the wire gauze between them (as shown in Fig. 18), and heat it vigorously until the crystals lose their color and turn to

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[1] To successfully make this experiment you must have a scale that will weigh down to a *centigram*.

a white powder. As soon as the crucible is cool enough to handle weigh it carefully again when you will find that it has lost weight.

This loss of weight is due to the water which was in the blue vitriol having been driven off and which was taken in at the time of crystallization; and this water is called the *water of crystallization*. To perform the experiment of making crystals from water put a little of the white powder from the crucible on a watch glass; let it stand for a few minutes when it will take back the water from the air and blue crystals of cupric sulphate will again be formed.

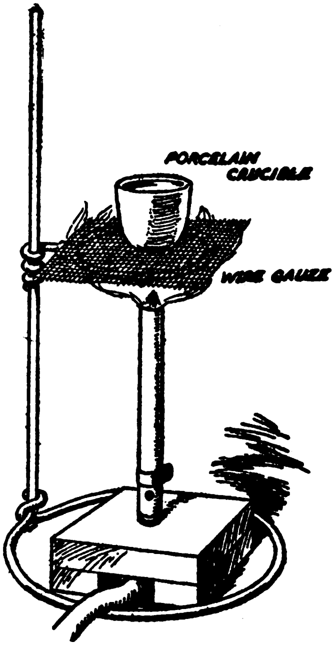

Fig. 18.—How to Heat a Crucible.

There are some hydrates that do not have to be heated to drive the water out of them but need only to be set in the open air when they lose the water in them and hence lose their crystallized forms. *Sodium sulphate* (Na_2SO_4) is one of these curious chemicals and such a substance is said to *effloresce* (pronounced *ef′-flo-resce* just as it is spelled).

Chemicals which Absorb Water.—Again there are other substances which when exposed to the air take on water. This process is called *deliquescence* (pronounced *del-i-ques′-ence*).

Calcium chloride ($CaCl_2 \cdot 2H_2O$) is a substance of this kind and it is largely used to absorb the water from gases when you want to get them perfectly dry; hence it is known as a *drying agent.*

How to Work Glass.—It is very often necessary to cut, bend and otherwise work glass in setting up chemical apparatus and these simple directions will make it easy for you to do so.

To Cut Glass Tubing.—To cut off a piece of glass tubing nick the glass at the point where you want it to break with the edge of your three-cornered file.

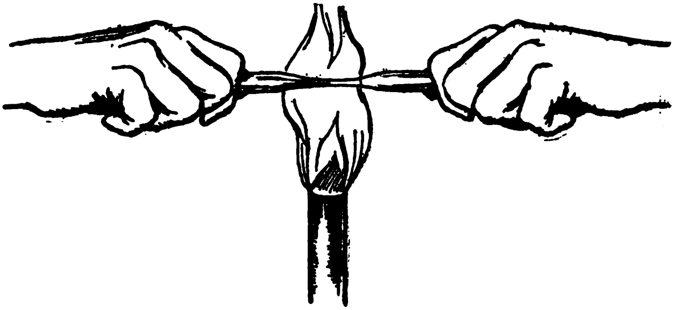

Fig. 19.—Drawing a Glass Nozzle.

Then hold the tube firmly in your fingers about 2 inches from both ends of the cut and apply enough pressure to break it. To smooth up the rough edges hold the ends of the tube in your Bunsen flame until they melt a little.

To Bend Glass Tubing.—This is an operation you should learn before you start to experiment. To bend a piece of glass tube hold it over the Bunsen flame until the tube is heated an inch or so on each side of the spot where you want to bend it.

Keep turning the tube until it is heated equally all around and becomes soft; then take it from the flame when you can bend it as you will if you do it gently.

After a very little practice you will be able to make a good smooth bend without the least trouble.

To Make a Glass Nozzle.—Many pieces of apparatus require a glass nozzle. To make one take a piece of glass tube 4 inches long, hold it at both ends over a Bunsen burner so that the flame strikes it in the middle, turn it slowly until the glass gets soft, stretch it by pulling gently on both ends at once until it is thin in the middle as shown in Fig. 19, then allow it to cool and cut it off in the middle; you will then have a nice little glass nozzle.

CHAPTER IV

ABOUT THE GAS CALLED AIR

- The Air Around the Earth
- What the Air is Made of
- The Lazy Gas of the Air
  - How to Make Nitrogen Gas
  - Some Experiments with Nitrogen
  - How Phosphorus Acts on Nitrogen
    - A Few Compounds of Phosphorus
  - The Physical Properties of Nitrogen
  - The Chemical Properties of Nitrogen
- The Heavy Gas of the Air
  - How to Make Carbon Dioxide
    - The Magic Glass of Water
    - The Cause Why of the Experiment
  - How to Make a Lot of Carbon Dioxide
    - An Astounding Experiment
  - The Physical Properties of Carbon Dioxide
  - The Chemical Properties of this Gas
- The Inert Gases of the Air
  - Argon, the Inactive Gas
  - Neon, Xenon, and Krypton Gases
  - Helium, the Sun Gas
- The Proof that Air is a Mixture
  - Liquid Air

CHAPTER IV

ABOUT THE GAS CALLED AIR

Next to water, air is about the most common thing you can think of and yet it is of the greatest importance to us in our daily lives—indeed, without air we could not live and there would be neither animal nor vegetable life on the earth.

The Air around the Earth.—Fortunately, either by accident or through design, there is a layer of air surrounding our earth that forms a covering about 50 miles thick and traces of air are known to exist at a height of 200 miles or so from the earth's surface.

If you were to make a balloon ascension and go straight up you would find that the air becomes thinner and thinner as you ascend and at a height of 6 or 8 miles you would not be able to breathe. (Ordinarily you draw into your lungs 500 c.c. or ½ a *liter* of air every breath you take.) The reason the air is thinner at the top than on the surface of the earth is because the weight of it above the surface pushes down on that which is below it and this makes the lower layers more dense than the upper layers.

It is in this thick part of the aerial ocean that animals —including ourselves—and plants are able to live. You perhaps know that if you climb a high mountain you will come to a place called the *timber line,* and trees and other vegetation will not grow beyond this line for the air is too thin and hence the temperature is too cold to support them. The idea is shown in Fig. 20.

Again without air, fire would be an unknown quantity and without fire we couldn't have the kind of pies that mother used to make or even the kind the baker bakes now, and so you can readily see that air is of the greatest importance to us.

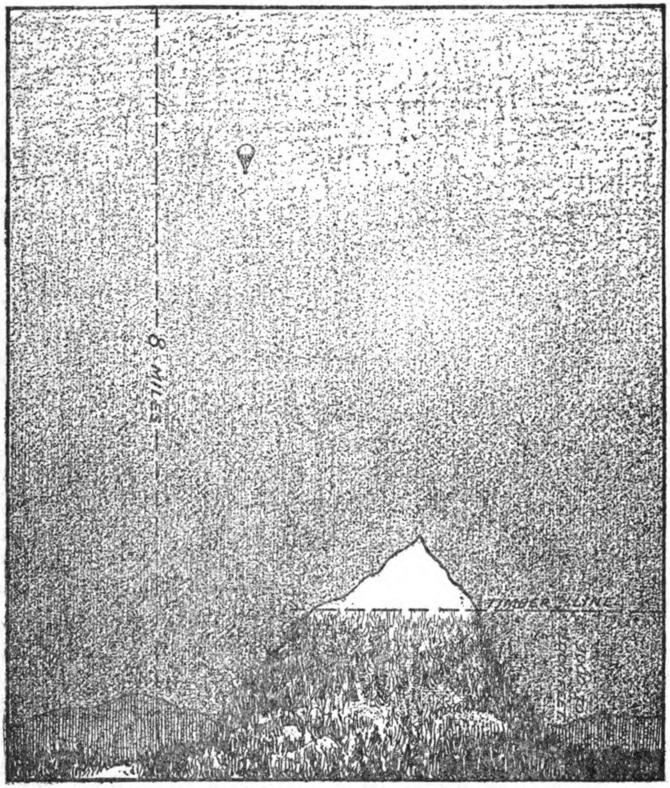

Fig. 20.—The Air above the Earth's Surface.

What the Air is Made of.—As I pointed out in the last chapter water is a chemical compound; now like water air is formed of two chief gases and these are *nitrogen* (N) and *oxygen* (O); but unlike water air is a mechanical mixture like sugar and sand, though it is not as easy to prove it. What I shall tell you about air a little farther on will show you, though, of the truth of it. While the two chief gases which form the air are nitrogen and oxygen it also contains six other gases plus considerable water vapor. Named in order of their importance these latter are *carbon dioxide, argon, neon, xenon* (pronounced *zē-non*) *krypton* and *helium.* Fig. 21 shows graphically the proportions of these gases in the air.

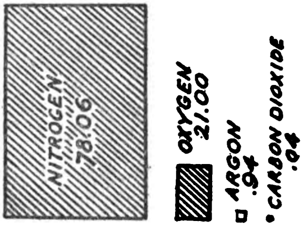

FIG. 21.—THE PROPORTIONS OF GASES IN THE AIR.

The Lazy Gas of the Air.—The nitrogen gas that is in the air makes up the most part of it for there are 78 parts by volume, or 75 parts by weight in a hundred parts of air. It is also found in combination with *potassium* (K) and *sodium* (Na) which comes from Chile and India (these compounds are called *potassium nitrate* (KNO_3) and *sodium nitrate* ($NaNO_3$) and also in the fat of animals.

HOW TO MAKE NITROGEN GAS.—You can easily make some nitrogen for yourself if you go about it like this: get a shallow dish, or pan, and fill it about half full of water; next look among your chemicals and get your *phosphorus* (P) bottle out.

It comes in small round white, or grayish, sticks about as big as your little finger and a couple of inches long. The bottle is filled with kerosene, or coal oil as it is more commonly called, and the sticks are kept in the oil for they burst into flame almost as soon as air strikes them **(WARNING—For this reason they are dangerous to handle unless you go about it right. More-**

over they are poisonous and you must keep your hands away from your mouth.)

So before you take the stick out of the bottle you want to have your apparatus all ready. Get your liter-wide-mouth bottle out and wipe it dry inside; now float your crucible in the dish of water and stand the liter bottle right beside the dish; next take the cap off of the phosphorus bottle and with your pliers pull a piece of phosphorus out and set it in the crucible.

Immediately set the liter bottle mouth downward over the crucible as shown in Fig. 22. The phosphorus will burst into flame and burn filling the bottle with a cloud of white smoke. Notice that the water in the dish rises in the bottle until it has about filled one-fifth of it.

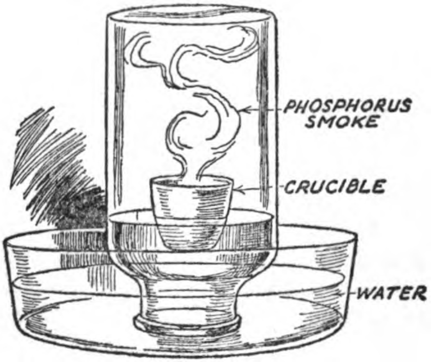

Fig. 22.—Making Nitrogen from the Air.

As soon as all the smoke has disappeared remove the crucible from the bottle taking care not to let any air enter the bottle and slip the piece of ground glass across the mouth of it. Hold it there firmly and remove the bottle from the dish, stand it upright on your bench and you will have a bottle full of nearly pure nitrogen.

Some Experiments with Nitrogen.—Hold a lighted match in the bottle and observe that the nitrogen does not support combustion, that is it does not burn for the match goes out.

Now capture a live fly and drop him into the bottle, replacing the ground glass top instantly. The fly will soon cease its struggles and die for just as nitrogen will not support combustion neither will it support life.

How Phosphorus Acts on Nitrogen.—As I told you the element that makes up the greater part of the air is nitrogen. Now although there are six other gases besides oxygen in the air still as there is only a trace of these gases present for all practical purposes we can consider the air as being composed of just the two gases, nitrogen and oxygen.

It must be clear now that if you can remove the oxygen from the air you will have only the nitrogen left and this is what the white phosphorus is for. When it burns it combines with the oxygen and in this way the latter is gotten rid of. Indeed had you burnt anything else in the crucible you would have got the same result, but I told you to use phosphorus because it burns easily of its own accord.

Further, when a substance burns it gives off gases, and unless you can find some way to remove these gases from the bottle the nitrogen will be impure and of course of no value for experimental work. Now when phosphorus burns in air the gases it gives off (this includes the smoke) dissolve very easily in water, and thus in the above experiment they remove themselves.

A Few Compounds of Phosphorus.—When the phosphorus burns there are two gases formed and these are known as *phosphoric*[1] *oxide* and *phosphorus*[1] *oxide.* The first is formed when the phosphorus starts to burn

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[1] The difference between the *ic* and *ous* endings on acids and salts is explained in Chapter VI. Also see Appendix II.

and there is plenty of air (that is oxygen) in the bottle for it to combine with.

The symbol for phosphorus is P and the symbol for nitrogen is N. Now in order to represent air and show that it is a mixture of nitrogen and oxygen and not a compound we will use a dash between the two symbols. Thus we will represent air as follows:

$$\overset{\text{Nitrogen}}{25N} - \overset{\text{Oxygen}}{5O}$$

showing that there is five times as much nitrogen in the air by volume as there is oxygen. This was proved by the fact that the water rose and filled just one-fifth of the bottle to take the place of the oxygen which had combined chemically with the phosphorus when it burned. This burning action can be shown thus:

$$\overset{\text{Phosphorus}}{2P} + \overset{\text{Air}}{25N}\text{—}\overset{\text{Oxygen}}{5O} \rightarrow \overset{\substack{\text{Phosphoric}\\ \text{oxide}}}{P_2O_5} + \overset{\text{Nitrogen}}{25N} \uparrow$$

(see footnote). Phosphoric oxide (or *phosphorus pentoxide* as it is sometimes called, because *pent* means 5) is the solid substance that is formed. It constantly dissolves in the water in the dish as explained in the chapter on acids.

As the oxygen is taken from the air, however, it grows less and still less until finally there is not enough to form phosphorus pentoxide. The phosphorus oxide (or *phosphorus trioxide*, *tri* meaning 3) is then formed for the formula for air has changed and is 25N—3O. The action now becomes

$$\overset{\text{Phosphorus}}{2P}+\overset{\text{Air}}{25N}\text{—}\overset{\text{Oxygen}}{3O} \rightarrow \overset{\substack{\text{Phosphorus}\\ \text{trioxide}}}{P_2O_3} + \overset{\text{Nitrogen}}{25N} \uparrow$$

This gas also dissolves immediately in the water.

Where an arrow points up it means that a gas is given off.

The Physical Properties of Nitrogen.—By *physical properties* I mean those things about a substance that you can *sense,* that is see, taste, smell or feel. If you will examine your bottle of nitrogen you will observe it is both colorless and odorless. It comprises ⅘ of the air by volume, oxygen making up the other fifth nearly. Since oxygen is slightly heavier than air it is natural that nitrogen should be slightly lighter than air.

Chemical Properties of Nitrogen.—Nitrogen does not, as you have observed, support either life or combustion as does oxygen; neither does it burn as does hydrogen. Although composing the greater part of the air it is useless as a life-giver.

It forms, however, many important compounds, among them being *gunpowder,* and nearly all other explosives, and in the more peaceful pursuits *fertilizers* for the soil and hence it is useful for these purposes. (See the chapters on *The Chemistry of Warfare,* and on *The Chemistry of the Soil.*)

As you saw in this experiment nitrogen does not combine or unite directly with phosphorus as does oxygen. Indeed nitrogen does not unite directly with any of the common elements as does oxygen (please bear in mind that all metals except gold, silver and platinum *oxidize,* that is rust).

For this reason nitrogen is called the *lazy element* while oxygen is called the *active element* of the air.

The Heavy Gas of the Air.—The next gas contained in air and which you must know about is *carbon dioxide.*

Properly speaking carbon dioxide does not form a part of pure air but since it is found in such large quantities in the layers of air nearest the surface of the earth it is an important element in it.

It is formed by the decay of vegetable matter, the breathing process, or *oxidization,* which is constantly going on in man and the lower animals, and the burning of all ordinary fuels such as wood, coal, and gas. It is also found in mines and is known by miners as the poisonous *choke-damp*.

Finally, it is found in many natural springs and the water from them is exactly the same as soda water, or *vichy*, which is water into which carbon dioxide has been forced. What is known as *hard water*, that is water with which soap will not form a lather, is caused by a mixture of carbon dioxide and *limestone* in the water.

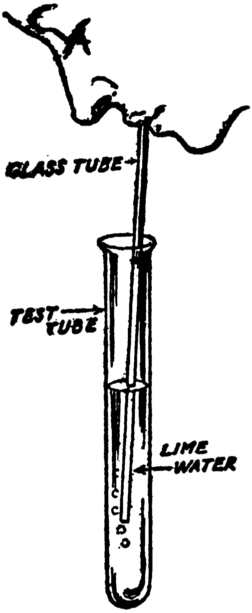

FIG. 23.—BLOWING CARBON DIOXIDE INTO LIME WATER.

How to Make Carbon Dioxide.—THE MAGIC GLASS OF WATER.—As mentioned before when animals breathe, carbon dioxide is formed and if you feel like a bear you can easily make some. To make a little of this gas take a test tube and half fill it with *calcium hydroxide*, the common name of which is *lime-water*.

You can make lime-water by dissolving a little ordinary mason's lime in water and shaking the test tube until the liquid becomes clear. Now put a short piece of glass tube down into the lime-water in the test tube. Next draw a long, deep breath into your lungs and hold it there for a few seconds; then put your lips to the other end of the glass tube and blow gently into the lime-water as shown in Fig. 23.

Slowly but surely it will turn a gray color and finally become milky-white. Stop blowing and let the test tube stand for a few minutes when you will find that a white powder has settled to the bottom of the tube leaving the liquid clear again. This is a *test* for determining the presence of *carbon dioxide* in a solution.

The Cause Why of the Experiment.—When you draw the deep breath the oxygen in the air combined with some of the used-up *tissues*[1] of your body which contain *carbon.*

This combination is again a case of slow burning, or *oxidization* as it is called. The symbol for carbon is C and the action which takes place inside of your body can be shown thus:

$$\overset{\text{of tissues}}{C} + \overset{\text{of air}}{2O} \rightarrow \overset{\text{Carbon dioxide}}{CO_2} + \text{heat}$$

You probably noticed that your breath was hot when you blew it out.

You will remember that you blew this carbon dioxide into lime-water that is a solution of calcium hydroxide. The symbol for *calcium* is Ca, and that of *hydroxide* is $Ca(OH)_2$.

A hydroxide is a substance which contains both *oxygen* and *hydrogen.* The parentheses is put around the OH in the formula to show that they are combined and act just the same as though they were one element instead of two separate and distinct elements. Such a combination is called a *radical.* This one is called the *hydroxide radical* (see *Base* in Chapter VI). As a matter of fact water is made of a hydroxide radical and hydrogen, thus:

$$\overset{\text{Hydrogen}}{H} + \overset{\text{Hydroxide}}{(OH)} \rightarrow \overset{\text{Water}}{H_2O}$$

But to return to the subject matter in hand. When you blew the carbon dioxide into the lime-water, the water acted on the carbon dioxide, and the two combined

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[1] These are formed of the cells of the body and their products.

to form another and single substance which is an acid known as *carbonic acid,* thus:

$$\overset{\text{Water}}{H_2O} + \overset{\substack{\text{Carbon}\\\text{dioxide}}}{CO_2} \rightarrow \overset{\substack{\text{Carbonic}\\\text{acid}}}{H_2CO_3}$$

There are many substances like carbon dioxide which act with water to form an acid. These substances when in a dry state are called the *anhydrides* which mean acids without water, of the acids they form with water. Thus dry carbon dioxide gas is the anhydride of carbonic acid. As usual you will notice that we have the same number of H's, C's and O's on both sides of the equation.[1]

Further when the carbonic acid was formed it reacted with the calcium hydroxide, to form calcium carbonate and water. Since calcium carbonate is insoluble, that is it will not dissolve in water this is the substance that gave to the liquid the milky-white appearance.

The action can be written thus:

$$\overset{\substack{\text{Carbonic}\\\text{acid}}}{H_2CO_3} + \overset{\substack{\text{Calcium}\\\text{hydroxide}}}{Ca(OH)_2} \rightarrow \overset{\substack{\text{Calcium}\\\text{carbonate}}}{CaCO_3} \downarrow + \overset{\text{Water}}{2H_2O}$$

Again you will observe that you have the same number of sybmols or letters on one side that you have on the other. That is to say, the *equation balances.*

When a number follows a radical in parentheses it multiplies the letters in the parentheses. Thus instead of writing $(OH)_2$ we could write it 2O2H but as this would be much longer in many cases it is easier to do it the other way.

The arrow after the $CaCO_3$ shows that the $CaCO_3$ does not dissolve in water. Such a substance is called a *precipitate,* and is always indicated by a downward

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[1] A proposition showing that two quantities are equal is called an *equation.*

arrow. A product which comes off as a gas is indicated by an upward arrow.

How to Make a Lot of Carbon Dioxide Gas.—Take a bottle or a Florence flask that is thoroughly dry, fit it with a stopper that has a hole in it and push a glass tube through it so that its lower end is close to the bottom as shown in Fig. 24.

Now place an equal amount of *cupric* (CuO) *oxide* that is powdered *copper oxide* and *wood charcoal* which you have mixed together in the bottle and tilt it a little. Set the bottle on the ring stand over your Bunsen burner and heat the contents of the bottle until it is red and for a few moments longer. You will then have a bottle full of *carbon dioxide gas.*

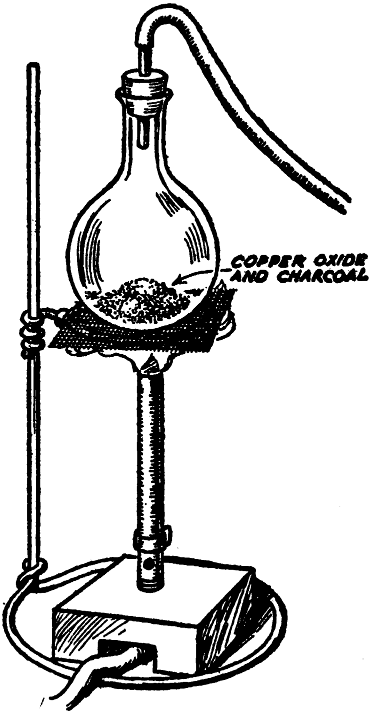

FIG. 24.—MAKING CARBON DIOXIDE ON A LARGER SCALE.

An Astounding Experiment.—If you now will set a dozen candles in a V-shaped trough with glass sides to it, as shown in Fig. 25, each candle will go out in succession as the carbon dioxide is poured out and onto the flames for the carbon dioxide displaces the oxygen and there is no longer enough of the latter gas present to support combustion.

THE PHYSICAL PROPERTIES OF CARBON DIOXIDE.—Carbon dioxide is an odorless, colorless, tasteless gas. It is one and a half times *heavier* than air, and with it some astonishing experiments can be performed as shown in Fig. 25. It is soluble in water and soda water is made by forcing the gas into the water under a pressure of from 60 to 150 pounds.

THE CHEMICAL PROPERTIES OF THIS GAS.—Carbon dioxide does not burn or support combustion or life.

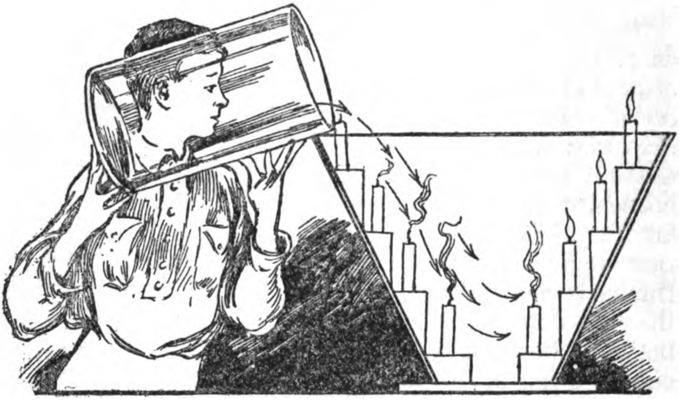

FIG. 25.—THE MAGIC CANDLE EXPERIMENT: POURING CARBON DIOXIDE DOWN ON TO LIGHTED CANDLES AND EXTINGUISHING THEM ONE BY ONE.

Indeed, as little as 4 per cent of it in a room will cause a flame to go out, while a little more will kill a human being.[1] From this you will see the necessity of ventilating rooms thoroughly and letting this gas out and fresh oxygen in. Since it does not support combustion it is like nitrogen, that is it is quite a lazy gas.

The Inert Gases of the Air.—Nearly all of the gases in the air were discovered by the fact that they unite

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[1] Not because it is a poison, but because it will not support life.

with other substances. There are however five other gases in the air known as the *inert*[1] *gases* because they do not form any known compounds; all of these gases were discovered in diverse, interesting and unexpected ways and so I'll tell you about them. These gases form altogether but 1 per cent of 100 per cent of the entire composition of the air but they are important because they are new elements.

ARGON, THE INACTIVE GAS.—In about 1892 Lord Rayleigh, an Englishman, was working with nitrogen and he discovered that the nitrogen from the air weighed a trifle more than that which he obtained from nitrogen compounds.

This fact lead him to believe that the nitrogen of the air must contain some other gas or gases which were much heavier. So he began to experiment and it was not long before he had separated an entirely unknown gas from the nitrogen.

He called this gas *argon*, which means inactive, because it would not enter into chemical combination with anything else. There is .94 per cent of argon by volume in the air; the other four gases which I shall tell you about make up the other .06 per cent thus making the total volume of the inert gases 1 per cent and the total volume of nitrogen and oxygen 99 per cent.

NEON, XENON, AND KRYPTON GASES.—Notwithstanding this new and epoch-making discovery Lord Rayleigh and Sir William Ramsay, anothe English chemist, again noticed a slight heaviness of the nitrogen which had been purified of argon as against that which was obtained from nitrogen compounds. The result of this was the discovery, after much experimenting, of three more new gases, for their proportion in nitrogen was exceeding small. They are *neon*, meaning *new*, *xenon* meaning *stranger*, and *krypton* meaning *hidden*.

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[1] Inert means inactive.

HELIUM, THE SUN GAS.—Finally Sir William Ramsay made the crowning discovery of *helium*, the fifth inert gas in the atmosphere. He called it helium from the Greek word *helio*, which means the sun, because it was first noticed on the sun by Lockyer in 1869,[1] who thought then that it was an element which we did not have here on earth.

The Proof that Air is a Mixture.—All during the course of this chapter I have told you many times that air is a mixture but nowhere have I actually proved this statement to you yet and no doubt you have been wondering how you can know it is a mixture and not a chemical compound.

LIQUID AIR.—You have probably heard of *liquid air* and it is by means of it that chemists are able to prove that air is really a mixture. Liquid air is air which has been subjected to such cold and pressure as to cause it to liquefy.

All of the components of ordinary gaseous air however are contained in it, and they are constantly struggling to get back into their gaseous state. Hence liquid air evaporates very rapidly.

Now when liquid air evaporates it has been found that the nitrogen gas passes off first and the oxygen next. This proves that air is a mechanical mixture and not a chemical compound for if it were the latter the nitrogen and oxygen would be combined so that they could not be separated in this way by evaporation since evaporation is merely a physical operation and not a chemical process.

[1] This is done by means of the spectroscope.

CHAPTER V

SOME OTHER COMMON GASES

- Ammonia, the Best Known Gas
    - How to Make Ammonia Gas
        - How the Experiments Work Out
    - The Physical Properties of Ammonia
    - The Chemical Properties of this Gas
- Hydrogen Sulphide, the Smelly Gas
- How to Make Hydrogen Sulphide
        - How the Experiments Work Out
    - The Physical Properties of Hydrogen Sulphide
    - The Chemical Properties of the Gas
- Sulphur Dioxide, the Heavy Gas
    - How to Make Sulphur Dioxide
        - An Explanation of the Experiments
    - The Physical Properties of Sulphur Dioxide
    - The Chemical Properties of this Gas
- Chlorine Gas, the Great Bleaching Agent
    - How to Make Chlorine Gas
        - An Explanation of the Experiment
    - The Physical Properties of Chlorine Gas
    - The Chemical Properties of the Gas

CHAPTER V

SOME OTHER COMMON GASES

WHILE you are experimenting with gases you might just as well learn about a few more of them that are generally useful. Named these are (1) *ammonia* gas, (2) *hydrogen sulphide* gas, (3) *sulphur dioxide* gas, and (4) *chlorine* gas.

These gases are often found useful in our home lives and some of them are employed on a large scale in industrial operations. You will have no difficulty in recognizing them, no, not even if you meet them in a blind alley on a dark night, for each of them is uncommonly smelly.

Ammonia, the Best Known Gas.—Where It Comes From.—Do not confuse *ammonia gas* with *aqua ammonia* sold by druggists and grocers. The latter is simply water in which a considerable amount of ammonia has been dissolved.

The air we breathe contains a goodly quantity of ammonia gas which has been formed by decaying vegetable matter. Ammonia gas is also given off when coal is distilled in making illuminating gas that is the kind of gas used in cities for lighting and heating. It is also given off when coke is made, and when oil is manufactured from shale, which is a kind of slate, large amounts of ammonia are formed.

HOW TO MAKE AMMONIA GAS.—To make ammonia gas you can use the same apparatus I described for making hydrogen. In the flask put a mixture of 10 grams each of *powdered quick-lime*, that is, unslaked lime, and *ammonia chloride* which you have mixed *intimately*[1] in your mortar.

Heat this mixture gently for a few seconds and then gingerly sniff around the end of the delivery tube and find whether or not you can smell any ammonia gas. If it has started to generate an odor similar to that of household ammonia will be present.

WARNING.—In the experiments described in this chapter do not get your nose too near the apparatus as the gases are strong and an overdose of them might prove injurious.

Now push the end of the delivery tube well up toward the bottom of a dry, inverted, wide-mouth bottle, as shown in Fig. 26, and still heating the flask gently wait until you smell a little ammonia in the air about you. Next draw the delivery tube from the bottle and instantly slip your piece of ground glass across the mouth of the bottle. Then put the end of the delivery tube in a

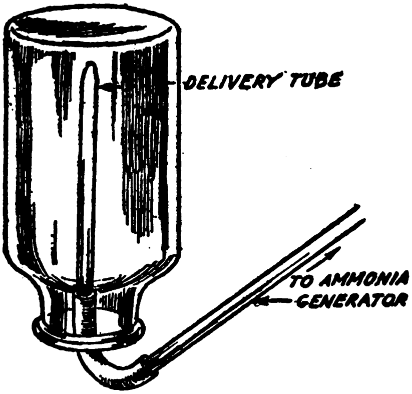

FIG. 26.—FILLING A WIDE-MOUTH BOTTLE WITH AMMONIA GAS.

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[1] This is the word that chemists use when they mean thoroughly.

small dish of water and let the ammonia gas bubble through it.

Examine the gas in the bottle and you will note that it is colorless but not odorless. Thrust a lighted splinter of wood into the bottle and observe that the gas does not burn and the flame of the splinter goes out. This done wet the palm of your hand, take off the ground glass and slip your hand firmly over the mouth of the bottle as shown in Fig. 27.

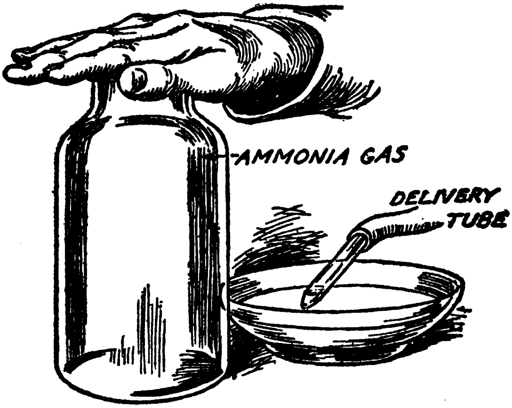

FIG. 27.—EXPERIMENT TO SHOW THE GREAT SOLUBILITY OF AMMONIA GAS IN WATER.

In a second or two try to remove your hand from the bottle and you will find that it seems as if it were being sucked into the mouth of it. Now examine the water in the dish and observe that while still colorless it smells strongly of ammonia gas; and this is *aqua ammonia.*

HOW THE EXPERIMENT WORKS OUT.—Slaked lime is, as I told you in the last chapter, known in chemistry as calcium hydroxide and its formula is $Ca(OH)_2$. Now ammonium chloride, the other ingredient used in making ammonia gas, is made up of three elements in combination and these are nitrogen, hydrogen, and chlorine. You will notice that this is ammon*ium* chloride and not ammon*ia* chloride.

Ammonium is a *radical*, that is it is a combination of two elements which always act as though they were one. The formula for the ammonium radical is NH_4. Hence the formula for ammonium chloride is NH_4Cl since it it is simply a compound of one part of ammonium and one part of chlorine.

When you mix the calcium hydroxide and the ammonium chloride and heat them a chemical *reaction* takes place between them. Since the action which takes place here is similar to many others that you will see later, take a good look at it.

When heat is applied the hydroxide or (OH) radical of the calcium hydroxide leaves it to combine with the ammonium radical of the ammonium chloride and a new compound known as *ammonium hydroxide* is formed. This leaves the chlorine and calcium free to combine with each other which they do, when they form *calcium chloride.*

Thus from your two original compounds you have gained two new ones. In other words the radicals of these compounds have simply exchanged places with one another. This happens in many actions in chemistry, and such a change of place between the radicals of two chemical compounds is called a *double decomposition.* It can be shown by a formula thus:

$$\underbrace{\overset{\substack{\text{Calcium}\\ \text{hydroxide}}}{Ca(OH)_2} + \overset{\substack{\text{Ammonium}\\ \text{chloride}}}{2NH_4Cl}} \rightarrow \overset{\substack{\text{Calcium}\\ \text{chloride}}}{CaCl_2} + \overset{\substack{\text{Ammonium}\\ \text{hydroxide}}}{2NH_4(OH)}$$

The lines show how these radicals change places to form new compounds.

Now when the ammonium hydroxide is still farther heated it breaks down and forms water and ammonia gas thus:

$$\overset{\text{Ammonium hydroxide}}{2NH_4OH} \rightarrow \overset{\text{Water}}{2H_2O} + \overset{\text{Ammonia}}{2NH_3} \uparrow$$

You will observe that the formula for ammonia is NH_3 while that of ammonium is NH_4. The arrow after the NH_3 shows that it is a gaseous product of the reaction.

THE PHYSICAL PROPERTIES OF AMMONIA.—As you have seen ammonia gas is colorless and has a pungent, sharp odor. It is twice as heavy as air and 600 c.c. of the gas will dissolve in 1 c.c. of water at room temperature and here follows the reason why when you moistened your hand it stuck to the bottle.

A large amount of the ammonia in the bottle dissolved in the water on your hand. This left a vacuum in the bottle and as a result the air pressing on your hand outside of the bottle forced your hand into it as far as it could go. When ammonia gas is put under pressure it easily becomes a liquid and this liquid is used in the manufacture of artificial ice. In evaporating the ammonia absorbs large quantities of heat from the water and thus causes the latter to freeze. The solution of ammonia water in the dish is the same as that used for household purposes of cleaning.

THE CHEMICAL PROPERTIES OF THIS GAS.—Ammonia neither supports combustion nor life nor does it burn in air. It does burn however in pure oxygen such as you made in Experiment 1, when it forms steam and nitrogen thus:

$$\overset{\text{Ammonia}}{2NH_3} + \overset{\text{Oxygen}}{3O} \rightarrow \overset{\text{Water}}{3H_2O} \uparrow + \overset{\text{Nitrogen}}{N} \uparrow$$

When ammonia is led into water, as you did when you put the delivery tube into the dish, a large amount of it is dissolved by the water and some of it act chemically with the water to form *ammonium hydroxide* which is a base. (See the chapter on *Acids, Bases and Salts.*)

It is this substance and not the dissolved ammonia, as many people seem to think, that cuts grease and makes household ammonia an effective cleaning agent. Its action is this:

$$\underset{\text{Ammonia}}{NH_3} + \underset{\text{Water}}{H_2O} \rightarrow \underset{\text{Ammonia hydroxide}}{NH_4OH}$$

Hydrogen Sulphide, the Smelly Gas.—When one of Mrs. Jones' boarders said he wanted "two boiled eggs, one of which *must* be good" he must have caught a fleeting odor of *hydrogen sulphide*[1] wafted across the breakfast table on the wings of a gentle morning zephyr.

This same vile-smelling gas is also present in some medicinal waters which are known as *sulphur waters* and these bubble forth from the earth in springs, which have their sources in the lower regions, and notwithstanding its odoriferous smell and obnoxious taste many people drink these waters for their health's sake.

HOW TO MAKE HYDROGEN SULPHIDE.—You can make some hydrogen sulphide gas with very little trouble if you go about it right. Using the hydrogen making apparatus again put 20 grams of *ferrous sulphide* in the flask.

Make a dilute solution of *hydrochloric acid* using 20 c.c. of water to 10 c.c. of acid, and pour this down the funnel onto the ferrous chloride. After a couple of seconds you will smell the hydrogen sulphide coming from the delivery tube. Wait a few seconds more and then light the gas issuing from it.

Note that it burns with a blue flame which can hardly be seen. Hold a cold dish in the flame as shown in Fig. 28, and you will see that a yellow deposit soon gathers

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[1] Hydrogen sulphide is the gas that is set up and liberated by rotten eggs when the shells are broken.

on the bottom of the dish, and that the gas in the flask is colorless.

How the Experiment Works Out.—The formula for ferrous sulphide is FeS and it is simply a compound of one part of iron to one part of sulphur.

When the hydrochloric acid is added to the ferrous sulphide a double decomposition takes place, that is, the sulphur combines with the hydrogen of the acid and this leaves the iron and chlorine free to combine which they do thus:

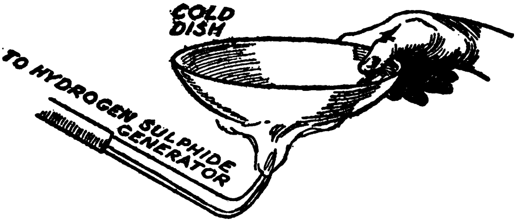

Fig. 28.—Collecting Sulphur formed by Burning Hydrogen Sulphide.

$$\overset{\text{Ferrous sulphide}}{FeS} + \overset{\text{Hydrochloric acid}}{2HCl} \rightarrow \overset{\text{Ferrous chloride}}{FeH_2} + \overset{\text{Hydrogen sulphide}}{H_2S} \uparrow$$

The iron and the chlorine combine to form a compound known as *ferrous chloride.*

Now when the hydrogen sulphide burns in the air it breaks down into water (or steam) and *sulphur dioxide* (a gas you will learn about later in this chapter), the reaction being like this:

$$\overset{\text{Hydrogen sulphide}}{H_2S} + \overset{\text{Oxygen}}{3O} \rightarrow \overset{\text{Water}}{H_2O} + \overset{\text{Sulphur dioxide}}{SO_2} \uparrow$$

However, the sulphur dioxide as soon as it is formed acts with the hydrogen sulphide coming from the delivery tube to form water again and free sulphur.

$$\overset{\text{Hydrogen sulphide}}{2H_2S} + \overset{\text{Sulphur dioxide}}{SO_2} \rightarrow \overset{\text{Water}}{2H_2O} + \overset{\text{Free sulphur}}{3S\downarrow}$$

(See footnote.) This free sulphur was the yellow deposit you saw on your dish.

THE PHYSICAL PROPERTIES OF HYDROGEN SULPHIDE.—As you have seen hydrogen sulphide is a colorless gas with an odor like unto that of ancient eggs which had failed to get into cold storage. Indeed, when an egg begins to go to the bad it is a regular hydrogen sulphide factory. This gas is lighter than air and three volumes of it will dissolve in one volume of water.

THE CHEMICAL PROPERTIES OF THE GAS.—When hydrogen sulphide burns in air it breaks down as described above. It dissolves in water and reacts to form an acid which is called *hydrosulphurous acid.* When it is led through a salt solution it precipitates *sulphides* of bright and varying colors.

Sulphur Dioxide, the Heavy Gas.—Sulphur dioxide is twice as heavy as air and it is largely used for *bleaching,* that is for removing the coloring matter from various materials.

This gas is found in large quantities near volcanoes and its odor is a familiar one to everybody for it is like that of a newly lit sulphur match; in other words it smells like the infernal regions—or at least the way our imaginations lead us to believe those regions smell.

HOW TO MAKE SULPHUR DIOXIDE GAS.—It will be a good scheme for you to make some sulphur dioxide so that you can examine it at first hand and test its bleaching properties for yourself. Rig up your hydrogen sul-

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Where an arrow points down it shows that the substance is precipitated.

phide generating apparatus as you did before and use the same chemicals. Take off the delivery tube and attach your U tube as shown in Fig. 29.

Now half fill the U tube with *concentrated sulphuric acid* and be careful not to get any on your hands or clothes. Next fasten the delivery tube to the U tube also as shown in Fig. 29 and start your apparatus to generating hydrogen sulphide. After a few seconds the hydrogen sulphide will drive all of the air out of the apparatus and you will be ready to then make sulphur dioxide.

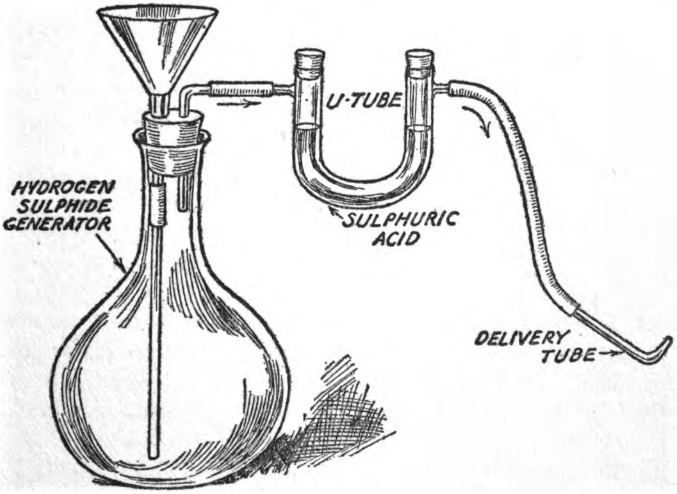

Fig. 29.—The Sulphur Dioxide Generator.

Place the end of the delivery tube in the bottom of a wide-mouth bottle and when you smell an odor similar to that which you smell when you strike a sulphur match remove the delivery tube and slip the ground-glass cover over the mouth of the bottle.

This done put the end of the delivery tube into a dish of water. To test its bleaching power slightly moisten a piece of colored silk with water, lift the cover from the jar, hang the silk down into the jar and put the cover back on it again as shown in Fig. 30. In the course of an hour or so the color will disappear entirely from the silk which shows in a striking manner that sulphur dioxide is a very good bleaching agent.

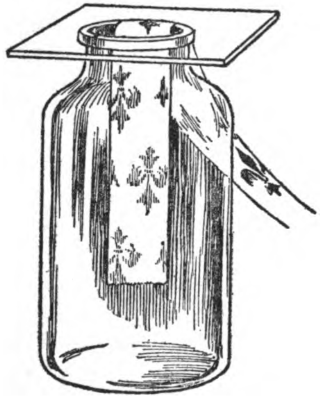

Fig. 30.—Bleaching Cloth in Sulphur Dioxide.

An experiment to show that this gas does not support combustion is to lower a lit match into the jar; you will see that the flame is immediately extinguished and that the gas does not burn. Also observe the yellow deposit that is formed in the bottom of the U tube.

An Explanation of the Experiments. — When the hydrogen sulphide bubbles through the sulphuric acid free sulphur is precipitated in the acid and water is formed. This accounts for the yellow deposit in the bottom of the U tube. The water mixes with the acid; sulphur dioxide[1] is also formed and this bubbles out of the acid and through the delivery tube. The action is as follows:

$$\underset{\text{Hydrogen sulphide}}{H_2S} + \underset{\text{Sulphuric acid}}{H_2SO_4} \rightarrow \underset{\text{Free sulphur}}{S\downarrow} + \underset{\text{Water}}{H_2O} + \underset{\text{Sulphur dioxide}}{SO_2\uparrow}$$

You will notice that the downward arrow in the equation shows the precipitate while the upward arrow shows the

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[1] When sulphur burns in air *sulphur dioxide* is formed, thus: $S + O_2 \rightarrow SO_2$. When sulphur burns in pure oxygen *sulphur trioxide* is formed, thus: $S + 3O \rightarrow SO_3$, and each of these gases has a different odor.

gas. The sulphur dioxide has the power of removing the delicate coloring matter from silks, straws and woolens.

THE PHYSICAL PROPERTIES OF SULPHUR DIOXIDE.—This gas is colorless and has a suffocating odor; as it is more than twice as heavy as air it can be poured from one vessel to another just like water. Fifty volumes of the gas will dissolve in one volume of water, hence you can see that it is quite soluble.

THE CHEMICAL PROPERTIES OF SULPHUR DIOXIDE.—Sulphur dioxide finds its chief uses as a disinfectant, that is it has the power to kill germs, and it is a bleaching agent. However, large quantities of the gas are manufactured for use in the sulphuric acid industry, which is explained in the next chapter.

Chlorine Gas, the Great Bleaching Agent.—And now we come to the last of these malodorous gases I have been telling you how to make and I think you will agree with me that they were all that I promised them to be. Like sulphur dioxide chlorine is largely used as a bleaching agent but unlike sulphur dioxide it does not occur free in nature.

HOW TO MAKE CHLORINE GAS.—You can easily make chlorine by using the apparatus described for breaking down water into its original constituents and which is shown in Fig. 17. When making chlorine, however, you will have to fill the dish and test tubes with a solution of 15 c.c. of hydrochloric acid in 50 c.c. of water.

The tube containing the chlorine will be filled with a greenish yellow gas and you will see the same amount of gas appears in each tube. When the tube is full of chlorine remove the carbon and fit a tight cork into the mouth being careful not to let any air enter.

Repeat this operation twice making two more tubes of hydrogen and corking each of them. Now moisten a piece of colored calico, which is made of cotton, and uncorking one of the tubes quickly drop the calico in, corking it again immediately. You will find that in a few minutes the goods will be bleached white.

Now stand the second tube in the test-tube rack; then uncork it and drop a pinch of finely powdered sodium into it and watch the brilliancy with which it burns, but keep your nose well away for chlorine is poisonous. Wet the ball of your thumb, uncork the third tube and clap your thumb over it when you will feel your thumb being sucked into the tube.

AN EXPLANATION OF THE EXPERIMENT.—Chlorine has the useful property of bleaching out dyes used for coloring fabrics, but as it is apt to hurt fine straws, silks and wools it is only used for cotton materials.

When having a good straw hat bleached you should always make sure that the hat cleaner is going to use sulphur dioxide which will not injure the straw and not chlorine. For an explanation of how electricity decomposes the hydrochloric acid see the Chapter on *Electrochemistry*. However, the current of electricity breaks it down thus:

$$\overset{\text{Sulphuric acid}}{\mathrm{HCl}} + \text{electricity} \rightarrow \overset{\text{Hydrogen}}{\mathrm{H}}\uparrow + \overset{\text{Chlorine}}{\mathrm{Cl}}\uparrow$$

Should you test the other tube with a lighted match you will get the test for hydrogen since this is what is generated and what the tube is filled with. The fact that the same amounts of hydrogen and chlorine are formed in this experiment proves that hydrochloric acid is made up of equal parts of hydrogen and chlorine.

The sodium, whose symbol is Na, burns in the chlorine, the two elements combining to form a cloud of white particles which are the same as common salt, or *sodium chloride* as it is called, thus:

$$\overset{\text{Sodium}}{\mathrm{Na}} + \overset{\text{Chlorine}}{\mathrm{Cl}} \rightarrow \overset{\text{Sodium chloride}}{\mathrm{NaCl}}$$

Your thumb stuck to the third tube for the reason explained under the experiment with ammonia.

The Physical Properties of Chlorine Gas.—Chlorine has a greenish yellow tint and a sharp penetrating odor. Indeed the word *chlorine* comes from the Greek and means pale green. The gas is two and a half times heavier than air and two volumes of it will dissolve in one volume of water at room temperature.

The Chemical Properties of the Gas.—Chlorine unites with many substances in the same way as it does with sodium. Among these substances are antimony and copper. When a jet of hydrogen burns in chlorine they form hydrogen chlorine gas, thus:

$$\overset{\text{Hydrogen}}{\mathrm{H}} + \overset{\text{Chlorine}}{\mathrm{Cl}} \rightarrow \overset{\substack{\text{Hydrogen}\\ \text{chloride}}}{\mathrm{HCl}}$$

When chlorine is led into water it dissolves and some HClO acid or *hypochlorous acid* is formed. It is this acid that really does the bleaching and not the chlorine itself. See Fig. 30. This is why it is necessary to wet the goods to be bleached. More information on bleaching will be found in the Chapter on *Chemistry In and Around the House.*

CHAPTER VI

ABOUT ACIDS, BASES AND SALTS

HOW ACIDS ACT

ACIDS YOU SHOULD KNOW ABOUT

WHAT AN ACID IS

HOW TO MAKE HYDROCHLORIC ACID

The Experiment Explained

HOW SULPHURIC ACID IS MADE

HOW TO MAKE NITRIC ACID

The Experiment Explained

THE CHEMICAL PROPERTIES OF ACIDS

ABOUT BASES, OR HYDROXIDES

SALTS AND HOW THEY ARE FORMED

KINDS OF CHEMICAL CHANGES

(*a*) Combination

(*b*) Decomposition

(*c*) Displacement, and

(*d*) Double Decomposition

REACTIONS OF ACIDS, BASES AND SALTS

CHAPTER VI

ABOUT ACIDS, BASES AND SALTS

You will recall that in the chapter on *water* I told you it is the greatest of all known solvents, which means that a larger number of chemicals will dissolve in water than in any other liquid. Thus water will dissolve nearly all of the *bases* and *salts* and a few of the lighter metals such as *potassium* and *sodium.*

How Acids Act.—There are, however, many elements, especially the heavier metals such as tin, lead, copper and iron that will not dissolve in water but which will dissolve very readily in acids.

Acids behave quite differently from water in that when other chemicals are dissolved the former usually acts on the latter and form new chemical compounds. It is this curious property of acids that makes them so valuable.

Acids You Should Know About.—There are quite a number of acids but there are only three that you should by all means know about for they are very useful and these are *hydrochloric*, *sulphuric* and *nitric acids.*

One or more of these acids is or are used in nearly every important industrial operation and they are therefore manufactured in large quantities. Indeed, it has been said that the civilization of a country can be measured by the amount of sulphuric acid it makes.

What an Acid Is.—Whenever a name is given to a class or a group of things it usually is done because each member of the class or group resembles each of the other members of it in one or more *specific* ways. For instance, when the name *acid* is applied to such different appearing liquids as hydrochloric, sulphuric and nitric acids, it is natural for us to try to find what they have in common with one another.

You have, more than likely, noticed that the formula for hydrochloric acid is HCl and that the formula for sulphuric acid is H_2SO_4 and I will tell you now that the formula for nitric acid is HNO_3. Now in the days when chemistry was young, it was thought that *oxygen* was the thing common to all acids, that is to say it was believed that all acids contained oxygen. In fact the very word *oxygen* means *acid former*.

Since then it has been found and you can plainly see, that oxygen is not the thing that the above acids contain but that it is *hydrogen*, and this is true of all acids. Hydrogen then is the real acid former, and hence in order for a compound to be an acid it must contain hydrogen.

Further, in order for an acid to be able to act on most substances it must first be dissolved in the greatest of all solvents, water. In fact water is as necessary to an acid as hydrogen is, for without water it will not show any of the properties which characterize an acid and which you will learn about a little later.

As you learned before a chemical compound which when dissolved in water forms an acid is known as the *anhydride* of that acid. But up until the time when it has been dissolved it is not an acid, for it has none of the properties of an acid. Thus *hydrogen chloride gas*, which I shall tell you how to make shortly, while it has the same chemical formula as hydrochloric acid is not an acid because it contains no water. But it is, however, the anhydride of hydrochloric acid.

And now I will tell you how to make two of the most useful acids for yourself and how a third one is made.

Of course the processes here described are different from those in commercial use, for these latter require large and expensive apparatus.

How to Make Hydrochloric Acid.—In the bottom of your hydrogen making apparatus, see Fig. 17, put about 20 grams of common salt. Fill a test tube nearly full of water and insert the end of the delivery tube into the test tube as shown in Fig. 31.

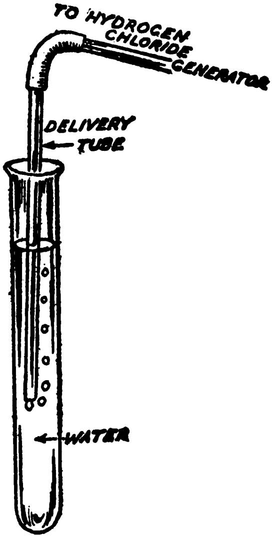

FIG. 31.—MAKING HYDROCHLORIC ACID.

This done pour a few drops of *concentrated* sulphuric acid into the funnel and give it time to trickle down onto the salt before you pour in any more. Between pours cover the top of the funnel well with the sheet of ground glass. Allow the gas which is generated to bubble through the water in the test tube for ten minutes and then fill the flask up with water. The contents of your test tube will then be *hydrochloric acid.*

If now you will dip a piece of *blue litmus paper*[1] into the acid you will see that it turns red. Next drop some small pieces of *granulated zinc* into the acid and observe that bubbles of gas are given off. Observe also that the acid is colorless, has a slight odor like that of chlorine and is a light, that is, thin liquid.

THE EXPERIMENT EXPLAINED.—When sulphuric acid is poured on common salt, or *sodium chloride,* it forms a compound called *sodium sulphate,* and *hydrogen chloride*

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[1] See Chapter IX.

gas is given off. The action is a *double decomposition* thus:

$$\overset{\substack{\text{Sodium}\\\text{chloride}}}{2NaCl} + \overset{\substack{\text{Sulphuric}\\\text{acid}}}{H_2SO_4} \rightarrow \overset{\substack{\text{Sodium}\\\text{sulphate}}}{Na_2SO_4} + \overset{\substack{\text{Hydrogen}\\\text{chloride}}}{2HCl} \uparrow$$

Now since 500 c.c. of hydrogen chloride gas will dissolve in 1 c.c. of water, the hydrogen chloride that passes into the test tube is immediately dissolved and the solution becomes *hydrochloric acid.*

How Sulphuric Acid is Made.—Sulphuric acid finds a far greater use in commerce than any other acid and hence its manufacture is carried out on a large scale. As it takes quite a thorough knowledge of chemistry, and quite a lot of expensive apparatus to make sulphuric acid successfully, I shall not tell you how to make it but rather how it is made.

The formula for sulphur*ic* acid is H_2SO_4. Now you will recall that the formula for sulphur dioxide is SO_2 and that when this gas is dissolved in water it forms sulphur*ous* acid, the formula of which is H_2SO_3 thus:

$$\overset{\substack{\text{Sulphur}\\\text{dioxide}}}{SO_2} + \overset{\text{Water}}{H_2O} \rightarrow \overset{\substack{\text{Sulphurous}\\\text{acid}}}{H_2SO_3}$$

You will notice that sulphur*ous* acid is similar in its make-up to sulphur*ic* acid except that it contains *one less* part of oxygen.

Kinds of Oxygen Acids.—There are two kinds of acids containing oxygen, namely (1) those in which the *least* oxygen is held in combination and these are known as *ous* acids, for example as H_2SO_3 sulphur*ous* acid and HNO_2 nitr*ous* acid; and (2) those in which the *most* oxygen is held in combination and these are known as *ic* acids, as H_2SO_4 sulphur*ic* acid and HNO_3 nitr*ic* acid.

Also it must be clear that if instead of dissolving sulphur *di*oxide in water you could dissolve sulphur *tri*oxide in water (the formula for which is SO_3) you would have sulphuric acid, thus:

$$\overset{\text{Sulphur trioxide}}{SO_3} + \overset{\text{Water}}{H_2O} \rightarrow \overset{\text{Sulphuric acid}}{H_2SO_4}$$

Here is one of the ways that sulphuric acid is manufactured commercially; first sulphur dioxide is passed over platinized asbestos with a stream of air and heated as shown in Fig. 32. The *platinized asbestos* acts just the same as the MnO_2 did in the oxygen experiment for it is a *catalytic agent.*[1] That is, it causes the sulphur dioxide to combine with the oxygen of the air more rapidly, but does not itself change. The sulphur dioxide combines with the oxygen of the air to form sulphur trioxide thus:

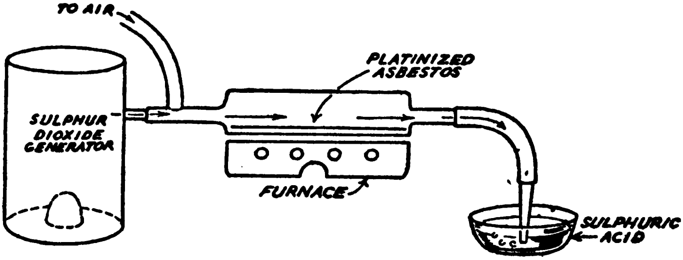

Fig. 32.—How Sulphuric Acid is Made.

$$\overset{\text{Sulphur dioxide}}{2SO_2} + \overset{\text{Of air}}{O_2} + \text{platinized asbestos} \rightarrow \overset{\text{Sulphur trioxide}}{2SO_3} + \text{platinized asbestos}$$

Now the sulphuric trioxide is a gas and does not dissolve easily in water but it dissolves very readily in concentrated sulphuric acid. Consequently instead of trying to make it dissolve in water a large amount of it is dissolved in a little concentrated sulphuric acid. Water

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[1] See Chapter III.

is then added to this solution when a large quantity of sulphuric acid is formed. This process is known as the *contact process* because of the catalytic, or contact agent used.

Pour a little sulphuric acid into a test tube and dilute it with water; dip a piece of blue litmus paper into the solution and you will see that it turns red. Next drop a few pieces of granulated zinc into the acid and note the bubbles of gas that are given off.

Also notice that the acid is oily and heavy and not thin and light like hydrochloric acid. The reason the chemists of the old school called it *oil of vitriol* was because it is of an oily nature and as it has a glassy appearance, the word *vitriol* when in the solid form was coined coming from the Latin *vitreous* meaning glassy in appearance.

How to Make Nitric Acid.—Powder 15 grams of *sodium nitrate* in your mortar and then put it in your 100 c.c. flask. Now rig up an arrangement like that shown in Fig. 33, first pouring just enough concentrated sulphuric acid upon the sodium nitrate to thoroughly moisten it.

Keep cold water running on the large flask and gently heat the small flask with your Bunsen burner. In order to keep a steady stream of water flowing on the flask you will find it convenient to attach a piece of rubber tube to the water faucet; you can then bring the end over the flask.

After a short time a yellowish brown liquid will begin to form on the sides of the flask and run down into the bottom and this is *nitric acid.* When you have made about a test tube full of the acid stop heating the flask and wait until no more gas is given off.

Next place a piece of blue litmus paper in the acid and again you will see that it turns red. Also drop a couple of pieces of copper slivers, or turnings, into the acid and note that a gas is given off. Also note the acid is a very light (mobile) liquid.

The Experiment Simply Explained.—When sodium nitrate is moistened with sulphuric acid, a sort of double decomposition takes place, giving *sodium bisulphate*, or *sodium acid sulphate* as it is sometimes called, and *nitric acid* thus:

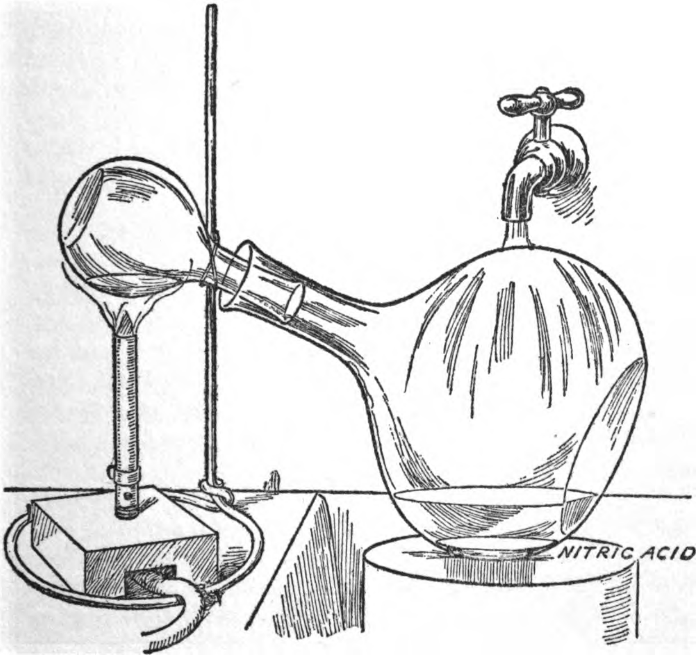

Fig. 33.—How to Make Nitric Acid.

$$\underset{\text{nitrate}}{\overset{\text{Sodium}}{NaNO_3}} + \underset{\text{acid}}{\overset{\text{Sulphuric}}{H_2SO_4}} \rightarrow \underset{\text{bisulphate}}{\overset{\text{Sodium}}{NaHSO_4}} + \underset{\text{acid}}{\overset{\text{Nitric}}{HNO_3}}$$

From the above formula you will observe that the sulphuric acid contains two parts of hydrogen. In ordinary double decomposition both of these parts would combine with the NO_3 and the other part remains in combination with the SO_4, to which the sodium joins itself.

Hence you have here a compound which is *apparently* a salt and yet which contains hydrogen, the acid former. It is for this reason that it is called *sodium acid sulphate.*

There are many other salts of this kind, that is, those that contain hydrogen, and these are known as *acid salts.*

The Chemical Properties of Acids.—As I told you before these three acids, namely, sulphuric, hydrochloric and nitric, are all called acids because they have certain properties in common with each other. For instance, from their formula you have seen that each one contains hydrogen. This hydrogen which is contained in every acid is known as the *acid radical* and it has the property of exchanging places with other radicals and elements very easily in double decompositions.

Now a number of metals, see Chapter VII, have the property of setting the hydrogen of an acid free. Thus when you dropped bits of zinc into hydrochloric acid the gas that you saw was hydrogen. At the same time that the hydrogen was given off a salt containing the metal and that part of the acid set free by the escaping hydrogen was formed, thus:

$$\begin{array}{ccccccc} \text{Hydrochloric acid} & & \text{Zinc} & & \text{Zinc cloride} & & \text{Hydrogen} \\ 2HCl & + & Zn & \rightarrow & \underset{\text{a salt}}{ZnCl} & + & 2H\uparrow \end{array}$$

Also when you dropped the zinc into the dilute sulphuric acid you saw bubbles of gas rise through the liquid and these were bubbles of hydrogen gas.

Another salt known as *zinc sulphate* was also formed, in this manner:

$$\overset{\text{Zinc}}{Zn} + \overset{\text{Sulphuric acid}}{H_2SO_4} \rightarrow \underset{\text{a salt}}{\overset{\text{Zinc sulphate}}{ZnSO_4}} + \overset{\text{Hydrogen}}{2H} \uparrow$$

When nitric acid is added to copper the metal displaces the hydrogen just as with the other acids, thus:

$$\overset{\text{Copper}}{Cu} + \overset{\text{Nitric acid}}{HNO_3} \rightarrow \overset{\text{Copper nitrate}}{Cu(NO_3)_2} + \overset{\text{Hydrogen}}{2H} \uparrow$$

However, as soon as the hydrogen forms, the nitric acid is so powerful that it acts with the hydrogen which is liberated and makes water and it also makes another gas called *nitric oxide,* thus:

$$\overset{\text{Nitric acid}}{NHO_3} + \overset{\text{Hydrogen}}{3H} \rightarrow \overset{\text{Water}}{2H_2O} + \overset{\text{Nitric oxide}}{NO} \uparrow$$

and so the gas that you saw coming off was not hydrogen in this case but nitric oxide, although hydrogen was at first set free. So you see nitric acid behaves just a little different from hydrochloric and sulphuric acids although its general action is the same.

And now let us sum up the chief things that are characteristic of the action of acids, to wit,

(1) When an acid is acted on by a metal the result is that a salt of the metal is always formed and that hydrogen is always set free.

(2) Furthermore you have noticed that all three *acids turned blue litmus paper red* and this is also a property which all acids have. Hence this is the recognized test for an acid.

About Bases, or Hydroxides.—In Chapter III I explained that a *base* was a compound in which (OH) or oxygen and hydrogen in combination is always found and that this (OH) is called the *basic* or *hydroxide radical.*

What I did not tell you though was that a base *always* contains the hydroxide radical in combination with a metal as in *sodium hydroxide* Na(OH), *calcium hydroxide* $Ca(OH)_2$, and *potassium hydroxide,* K(OH), in which sodium, calcium, and potassium are all metals. Now, let us see how these bases, or hydroxides are formed.

How Bases are Formed.—When water is added to the oxides of the above metals a violent action takes place, and in the case of calcium oxide, great clouds of steam are produced. These reactions are as follows:

$$\overset{\text{Sodium oxide}}{Na_2O} + \overset{\text{Water}}{H_2O} \rightarrow \overset{\text{Sodium hydroxide}}{2NaOH}$$

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

$$\overset{\text{Potassium oxide}}{K_2O} + \overset{\text{Water}}{H_2O} \rightarrow \overset{\text{Potassium hydroxide}}{2KOH}.$$

But sodium, calcium and potassium are very active metals and hence the ease with which they unite with water. With others of the less active metals such as lead, tin, and iron, although they do act slightly with water, the action is so slow that their hydroxides must be prepared in some other way. But this need not concern you now.

Take your sodium hydroxide solution from your shelf of chemicals, dip a piece of *blue* litmus paper into it and notice that it does not change color. Dip a piece of *red* litmus paper in it and observe that it *turns blue in the base.* This is the regular test for a hydroxide.

An interesting little experiment is to dip a piece of blue litmus into an acid when it turns red, and then dip it into a base when it turns blue again.

When an acid acts on any base, water is always formed together with a salt, thus:

$$\overset{\text{Hydrochloric acid}}{HCl} + \overset{\text{Sodium hydroxide}}{Na(OH)} \rightarrow \overset{\text{Sodium chloride}}{NaCl} + \overset{\text{Water}}{H_2O}.$$

This is a double decomposition but it goes by the name of *neutralization* because the solution no longer shows any of the properties of either acid or base, that is to say it is *neutral.*

What Neutralization Is.—The process of neutralization of an acid by a base can be strikingly shown by means of an *indicator* known as *phenolphthalein*[1] (pronounced *pen-ol′-ta-lean*). This is a colorless substance one drop of which when added to a solution of a base causes the solution to turn a deep pink.

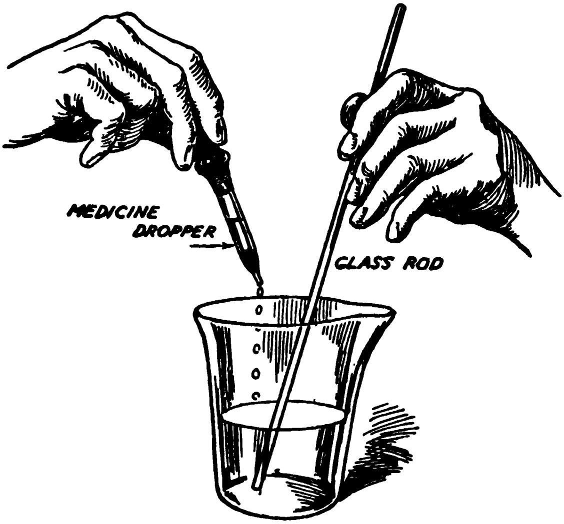

FIG. 34.—NEUTRALIZING AN ACID WITH A BASE.

Measure out 10 cubic centimeters of any kind of dilute acid in your graduated cylinder and pour it into a beaker. Wash the cylinder clean, and then measure out about 50 c.c. of a strong, that is, concentrated base,

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[1] See Chapter IX.

say sodium hydroxide solution and pour this into another beaker. Now add one drop of phenolphthalein solution to the beaker containing the solution.

Fill a medicine dropper, or fountain pen filler, with the base from the other beaker and drop by drop add it to the beaker containing the acid, and stir it well between each drop, as shown in Fig. 34. Finally the acid will just be neutralized and as you add the next drop of base the excess will cause the solution to turn a pale pink due to the phenolphthalein.

Now pour back into the graduated cylinder any unused base that remains in the beaker, and empty the rest of the base in the medicine dropper into it. Take a *reading*[1] of the amount of the base in the cylinder and subtract it from 50 c.c. the amount you started with. The result will be the number of cubic centimeters of base required to neutralize 10 c.c. of acid. This process is known as *titration*.

Salts and How They are Formed.—As you have already seen *salts* are formed by the action of metals on acids. However, we can go still farther and make the broad statement that every metal can form a salt with every acid.

Now the action:

$$\overset{\text{Zinc}}{\mathrm{Zn}} + \overset{\substack{\text{Hydrochloric}\\ \text{acid}}}{\mathrm{2HCl}} \rightarrow \overset{\substack{\text{Zinc}\\ \text{chloride}}}{\mathrm{ZnCl_2}} + \overset{\text{Hydrogen}}{\mathrm{2H}} \uparrow$$

is called a *displacement* because the hydrogen takes the place of, that is, it is displaced, from the acid by the metal. The action of all metals on acids to form salts are displacements.

Kinds of Chemical Changes.—Thus far you have met with four kinds of chemical changes, namely, (1) *com-*

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[1] Which means to note the amount.

bination, (2) *decomposition*, (3) *displacement*, and (4) *double decomposition.*

(1) *Combination.* To illustrate combination take the experiment with mercury and sulphur

$$Hg+S \rightarrow HgS.$$

Here two elements have combined to form one compound.

(2) *Decomposition.*—The oxygen experiment shows how decomposition takes place, thus:

$$\overset{\text{Potassium chlorate}}{KClO_3} \rightarrow \overset{\text{Potassium chloride}}{KCl} + \overset{\text{Oxygen}}{3O}.$$

In this case a compound is *decomposed* or broken down into its original elements.

(3) *Displacement.*—Where displacement takes place the hydrogen is displaced by a metal as described on page 94, and

(4) *Double Decomposition.*—In this reaction two compounds are decomposed and recombine in a different order from the original, thus:

$$H_2SO_4+2NaCl \rightarrow Na_2SO_4+2HCl\uparrow.$$

These are the four chief kinds of chemical actions and you should study every action you make to see which one of these actions it is.

Reactions of Acids, Bases and Salts.—And now that you know the four different kinds of actions in general you should learn those that occur between *acids*, *bases* and *salts* in particular. There are four of these actions and to keep them in mind will give you an everlasting end neverfailing key to almost all chemical phenomena.

(1) An *acid* and a *base* give *water* and a *salt*. (This is neutralization.)

(2) An *acid* and a *salt* give a *salt* and an *acid*. (This is a double decomposition.)

(3) *Two salts* give *two salts*. (This is also double decomposition.)

(4) A *base* and a *salt* give a *salt* and a *base*. (This is likewise double decomposition.)

In all cases but the first you will observe that the reactions are double decompositions and this will also help you when you are trying to discover what takes place in reactions between acids, bases and salts.

CHAPTER VII

METALS, ALLOYS AND AMALGAMS

- THE MOST USEFUL METAL, IRON
  - A Common Iron Ore
- THE METAL OF THE ANCIENTS, COPPER
  - Some of the Uses of Copper
- THE LIQUID METAL, MERCURY
- THE ELECTRIC METAL, ZINC
- THE SOFT, WHITE METAL, TIN
- THE HEAVY METAL, LEAD
- THE LIGHT METAL, ALUMINUM
  - Aluminum in Precious Stones
  - The Thermit Process of Making Pure Metals
  - The Thermit Process of Welding Metals
- THE METAL THAT BURNS, MAGNESIUM
- THE HARDENING METAL, MANGANESE
- THE COLOR METAL, CHROMIUM
- NICKEL, THE METAL THAT WILL NOT RUST
- THE EXPANDING METAL, ANTIMONY
- THE FUSIBLE METAL, BISMUTH
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- THE PRECIOUS METALS, SILVER, GOLD, AND PLATINUM
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  - Alloys of Iron and Steel
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CHAPTER VII

METALS, ALLOYS AND AMALGAMS

Now that you have a pretty good idea about acids, bases and salts you must know something about the metals. There are many other metals besides the ordinary ones, such as iron, copper, tin, lead, silver and gold and they are all highly interesting and useful.

The Most Useful Metal, Iron.—After the *stone age* came the *bronze age* and following this came the *iron age;* these three ages mark roughly the prehistoric stages through which the human race progressed from savagery to civilization.

Iron (Fe) is the most common and the most useful of all the metals and it is mined from the bowels of the earth. But it is not found in a pure state and so cannot be made immediately into knives, stoves, implements, machinery, etc.

Instead it is usually found in *ores*, that is in combination with one or more other elements and from which it must be separated. There are several ways by which this can be done but the process of extraction depends entirely on the kind of a compound that the iron is associated with.

A COMMON IRON ORE.—The most common of these compounds, or ores, as they are called, is *iron sulphide* or *ferrous sulphide* to give it its right name, the formula of which is FeS. In order to get the free iron from the ore it is *roasted*, that is burnt, in a large furnace with air blowing on it as shown in Fig. 35.

The oxygen of the air then combines with the sulphur and leaves the free iron, thus:

$$\overset{\substack{\text{Ferrous}\\\text{sulphide}}}{FeS} + \overset{\substack{\text{Of}\\\text{air}}}{O_2} \rightarrow \overset{\substack{\text{Free}\\\text{iron}}}{Fe} + \overset{\substack{\text{Sulphur}\\\text{dioxide}}}{SO_2}\uparrow .$$

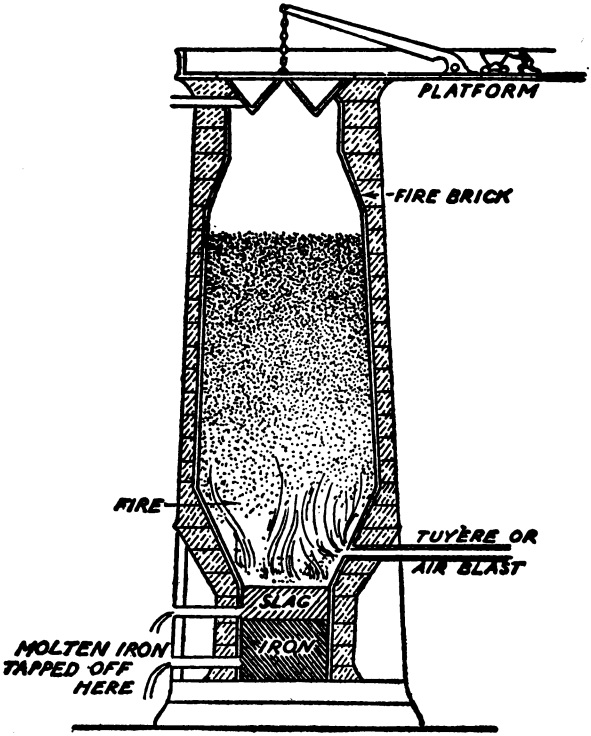

Fig. 35.—How Iron Ores are Reduced in a Blast Furnace.

Thus sulphur dioxide is formed and passes off as a gas leaving the iron free. This is a common process for extracting many metals that are found in combination with other elements, and which can be burnt off as stated above.

WHAT IRON RUST IS.—When iron stands in moist air the oxygen of the air slowly combines with it and forms *ferric oxide,* or *rust,* as it is called, the formula of which is Fe_2O_3 thus:

$$\underset{\text{Iron}}{2Fe} + \underset{\text{Oxygen}}{3O} \rightarrow \underset{\text{Ferric oxide}}{Fe_2O_3}.$$

If now you leave iron in moist air long enough it will rust away almost entirely. This is because the ferric oxide is of a loose, scaly nature and does not cover the iron tightly enough to prevent the oxygen of the air from attacking the iron still farther. Other metals, such as zinc, rust hardly at all for the rust coating clings to the surface of the metal so tight that the oxygen of the air cannot pass through it and act farther on it.

When iron is added to hydrochloric acid, or to any other strong acid, it sets the hydrogen of the acid free and an *iron salt* is formed thus:

$$\underset{\text{Iron}}{Fe} + \underset{\text{Hydrochloric acid}}{2HCl} \rightarrow \underset{\text{Ferrous chloride}}{FeCl_2} + \underset{\text{Free hydrogen}}{2H} \uparrow.$$

This salt, or compound, which is $FeCl_2$, is known as *ferrous chloride.*

The Metal of the Ancients, Copper.—Copper (Cu) was the first metal to be used by primitive man for making tools and utensils because it is found free in nature, and consequently did not require roasting or any other treatment to be of immediate use.

SOME OF THE USES OF COPPER.—As you have known ever since you were old enough to ask for one, copper is used for making those coins of trifling value called the *penny.*

I have mentioned before that copper does not set hydrogen free from nitric acid and since copper is not acted on by sulphuric acid it is used for making pumps where these acids, or compounds containing them have to be used.

It is also employed in making shells for rifles and guns because it does not rust and is very strong. Gunmetal and German silver are alloys of copper, that is to say they are mixtures of copper and other metals, melted together.

The Liquid Metal, Mercury.—This was the favorite metal of the alchemists and with it they worked their wonders, wondering as they worked. *Mercury* (Hg) is different from all other metals in that it is a liquid at ordinary temperatures.

It is found in little drops mixed with earthy substances and with *cinnabar* (HgS) which is *mercuric sulphide*. To get the free mercury the cinnabar is roasted when the sulphur is changed into the dioxide and the mercury is vaporized and then condensed, thus:

$$\overset{\substack{\text{Mercuric}\\ \text{sulphide}}}{HgS} + \overset{\text{Oxygen}}{O_2} \rightarrow \overset{\substack{\text{Metallic}\\ \text{mercury}}}{Hg} + \overset{\substack{\text{Sulphur}\\ \text{dioxide}}}{SO_2}.\uparrow$$

It is largely used in making amalgams which you will read about farther on.

The Electric Metal, Zinc.—I have called zinc (Zn) the *electric metal* in virtue of the fact that it is the most active of all the metals in generating current electricity when it is used as the negative element in a battery cell.

Zinc, like iron, is found in combination with sulphur in the form of *zinc sulphide* which must be roasted before it can be obtained as free zinc and the process is the same as that used to extract iron from its ores. (See Fig. 36.)

Zinc rusts a little, but it soon becomes covered thick enough so that it cannot rust farther. As you learned in Chapter II zinc sets the hydrogen from hydrochloric acid free and leaves *zinc chloride* and when this is dissolved in water it makes a good *soldering fluid.* Zinc is largely used for lining sinks and tanks, for roofing and wherever a metal must be used that is exposed to either air, water or both.

The Soft White Metal, Tin.—In the days of Herodotus, the ancient Greek historian, tin (Sn) was only found in the British Isles, and hence these early peoples called them the *Tin Islands.*

Tin is usually found in an oxide called *tinstone,* that is the tin is combined with oxygen and must be roasted to get it pure. However, it does occur free once in a while as a rare mineral. Tin does not rust at all, and is, therefore, used for making *tin-plate,* which is thin sheet iron coated with tin. Tin-plate in turn is largely used for roofing and culinary utensils.

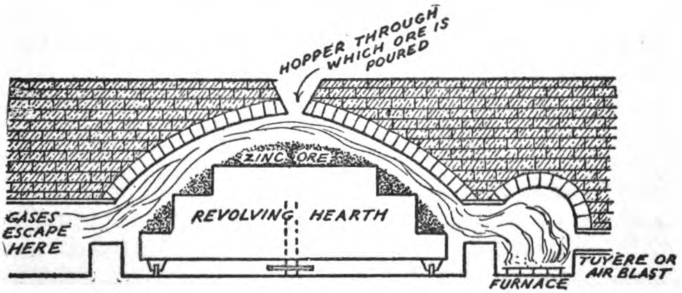

FIG. 36.—HOW ZINC ORE IS REDUCED.

Tin sets hydrogen free from *warm* hydrochloric acid but not from it when it is cold as iron and zinc do. Pure tin has the peculiar property of turning into a gray powder when subjected to a very cold temperature for a long time. This happened to a shipment of tin stored in the custom house at Petrograd, much to the disgust of its owners, and the amazement of the officials.

The Heavy Metal, Lead.—Lead (Pb) is an old time metal and if you will read the *Book of Job* you will find it mentioned there. It is said to have been used to glaze earthenware in the time of Solomon while it is known to a certainty that the Romans used it for weights, utensils and waterpipes.

Now while a pound of lead and a pound of feathers weigh the same, still lead is much the heavier bulk for bulk. Like iron it is found in combination with sulphur and to obtain it pure it has to be roasted. It is, however, sometimes found free though never in large quantities.

Lead does not rust, and it displaces the hydrogen from hydrochloric acid very slowly. Because it acts so very slowly with acids it has many uses in commerce, such as lining chambers and vessels in which sulphuric acid is kept.

It is also largely used in making shot, shrapnel bullets, and many other articles. Solder is made of 50 per cent of lead and 50 per cent of tin, and when the two are combined the alloy thus formed melts at a much lower temperature than either the lead or tin alone. For this reason lead or tin articles can be soldered without themselves being melted.

Because it is so easy to work and solder it is much used for plumbing pipes, especially where *hard water* is had. The salts in the water which make it hard form a protective coating over the surface of the lead and this prevents it from forming compounds which might otherwise injure the water.

Rain water does not contain these salts and is known as *soft water;* it does contain free oxygen in solution, though, and this will act slowly on a lead pipe to form *lead hydro-oxide* ($Pb(OH)_2$). For this reason tin and not lead pipes should be used for carrying soft water for drinking purposes.

The Light Metal, Aluminum.—The wonderful metal we call *aluminum* (Al) though quite recently discovered is one of the most widely distributed elements in nature. It was first obtained by treating *aluminum chloride* with *metallic sodium* but it is only made now on a large scale by *electrolysis,* shown in Fig. 37. It is a process invented in 1886 by C. M. Hall.

Aluminum is the lightest common metal known. It is found in combination with oxygen, *ortho*[1]*-silicates* of calcium and phosphates. Clay and rocks are chiefly composed of *aluminum silicates* and pure *aluminum* is obtained by decomposing them by an electric current.

ALUMINUM IN PRECIOUS STONES.—The semi-precious stones we call the *garnet* is formed of aluminum and calcium orthosilicate. The *torquoise*, another semi-precious stone, is composed of *aluminum phosphate*.

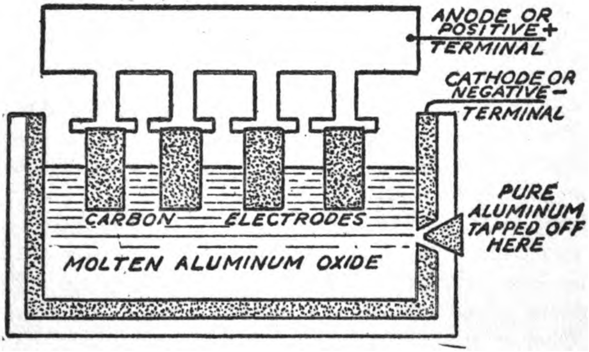

FIG. 37.—HOW ALUMINUM OXIDE IS ELECTROLYZED AND ALUMINUM IS OBTAINED.

The *ruby*, which is more costly than the diamond, is made up of *aluminum oxide* heated to a certain temperature and colored red with *chromium;* while the *sapphire*, another precious stone, is formed of *aluminum oxide* and colored blue with *aluminate of iron*. All of these gems can be made *synthetically*, that is treating the substances named above in an oxy-hydrogen furnace. (See Chapter XV.)

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[1] In chemistry, *ortho*, from the Greek *orthos*, which means *right*, is used to show those substances which contain the largest possible number of *hydroxyl molecules*.

Emery, which is used for making grinding wheels, is also made of aluminum oxide. Because aluminum is so light and strong it is largely used in making parts of airplane engines and other machinery. Aluminum sets the hydrogen of acids free and it rusts slightly but as the rust coats the surface closely it keeps out the air and so it cannot penetrate very deep.

One of the most valuable properties of aluminum is its great affinity for oxygen for it combines with all the oxides of the metals below *magnesium* in activity, a list of which you will find farther on in this chapter.

The Thermit[1] Process of Making Pure Metals.—Goldschmidt, a chemist, who was aware of the above fact made use of it and got up a process for the reduction of the oxides of metals—which in this case means that the oxygen of the metals are taken away from them—and this not only makes it possible to obtain pure metals but also to weld them, which while simple in the extreme is of great value.

In the Goldschmidt *thermit process*, as it is called, a mixture of powdered aluminum is put into a crucible together with powdered *ferric oxide* and the whole mass is then fired by means of a piece of burning *magnesium* when the aluminum will combine with the oxygen of the ferric oxide and this sets the iron free.

A temperature of 3000 degrees Centigrade melts both the iron and the *aluminum oxide* which, since they will not mix, separate into two layers, thus:

$$\overset{\text{Ferric oxide}}{Fe_2O_3} + \overset{\text{Aluminum}}{2Al} \rightarrow \overset{\text{Aliminum oxide}}{Al_2O_3} + \overset{\text{Iron}}{2Fe}.$$

The Thermit Process of Welding.—The very high temperature which is set up by the *thermit* makes it pos-

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[1] A trade name taken from the Greek word *thermo* which means hot.

sible to place the mixture of powdered aluminum and ferric oxide around the broken parts of an iron or a steel casting and ignite it, when the melted iron will flow between the broken parts, melt the edges of them, fill up the crack between them and so weld them securely together.

From what I have told you, and it is only a part of the story of this metal, it is evident that aluminum is a very useful element.

The Metal that Burns, Magnesium.—This metal is very light and brittle and it is used chiefly for producing a bright, sunlike light. It can be bought in either powder or ribbon form and as it can be lit with a match it is largely used for making photographs, fireworks and signaling.

Magnesium is not found free in nature but is obtained by the *electrolysis* [1] of melted *magnesium chloride* ($MgCl_2$). It is silvery white in its pure form and is very soft; as a metal it is active and will displace hydrogen from cold acids very rapidly.

It finds its chief use in the manufacture of flashlight powder for it has the property of giving off a brilliant white light when it burns, of such short wave lengths that it acts on a photographic plate just as sunlight does. Flashlight powder consists of 10 parts of powdered magnesium mixed with 17 parts of potassium chlorate.

The Hardening Metal, Manganese.—One of the most important uses of this metal is to harden steel. *Manganese* (Mn) steel is, of course, an alloy of manganese and steel and is largely used for railway *frogs* and rails, for safes, armor plates for ships, parts of heavy machinery and for bearings for automobiles.

Manganese is a grayish colored metal and is found only in compounds, the chief one of which is *manganese dioxide*, or *pyrolusite* (MnO_2) as it is called. This compound is

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[1] *Electrolysis* is a process of decomposing a chemical compound by passing an electric current through it.

easily reduced and the pure metal obtained by the *thermit process*, the MnO_2 taking the place of the Fe_2O_3 in the action. Manganese is of great value because of the *alloys* it forms with iron and about which more will be said presently.

The Color Metal, Chromium.—This curious metal gets its name from the Greek word *chroma* which means surface color. It does not rust but when heated it burns in oxygen and gives *green chromic oxide* (Cr_2O_3).

Chromium (Cr) is a white, crystalline metal that is very hard and, like manganese, it is prepared from Cr_2O_3 by the *thermit* process and, like magnanese, when small amounts are added to steel it makes it very hard and tenacious.

Nickel, the Metal that Will Not Rust.—Nickel is a white, hard metal which takes a high polish. It is never found free in nature, except in *meteors*,[1] but it is obtained generally from *pentanolite* which is a mixture of nickel, copper, iron and sulphur. This mixture is first roasted and then *smelted*[2] during its production. Its best known use is in making the *nickel*, that is our 5-cent piece, and for coating other metals, especially iron, by a process of electrolysis, commonly called *electroplating*. (See Chapter XI.)

The Expanding Metal, Antimony.—*Antimony* (Sb) is a beautiful, white lustrous metal. It is obtained by a reduction of *antimony oxide* (SbO_4).

It has the property of expanding when cooled after it has been melted and this makes it valuable for mixing with lead, which contracts, for casting type whose edges

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[1] A meteor is a small piece of a comet that has broken up and which gets within range of the earth's gravitational force when it is pulled to its surface. Many meteors contain nickel in its free state.

[2] Smelting means a process for obtaining a metal by fusion, that is heating it.

must be very sharp. It is also mixed with lead and tin to form *Babbitt metal* and other *anti-friction* bearings for machinery.

The Fusible Metal, Bismuth.—*Bismuth* (Bi) is a metal found free in nature and is quite like antimony. When mixed with lead it forms an alloy called *Wood's metal,* the melting-point of which is lower than the boiling-point of water.

This makes it possible to use the alloy for *electric fuses,* in fire-alarm systems and for *safety plugs* in boilers and in *automatic sprinklers* in buildings as shown in Fig. 38. When a fire breaks out and the heat gets up to 50.5 degrees Centigrade the plug of Wood's metal in the nozzle melts and this lets the water flow when it strikes a cap which sprays it in all directions.

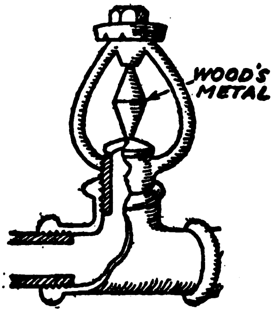


Fig. 38.—A Fire Sprinkler.

The Precious Metals, Silver, Gold and Platinum.—Of these precious metals silver (Ag) is the least, gold (Au) is the next, and *platinum* (Pt) is the most precious. All of them are found free in nature in California, Nevada, Mexico and Australia and these are obtained by what is called *placer mining* as shown in Fig. 39.

The striving of man to wrest these metals from the earth, the mining camps they built, and the blood they shed has furnished some of the most romantic scenes for which the Golden West is noted.

None of these metals will set hydrogen free from acids and none of them rust. Neither gold nor platinum will dissolve in either hydrochloric, nitric or sulphuric acid, but they will dissolve in hydrochloric and nitric acid when they are mixed together. The early chemists called this solution *aqua regia* because it alone would dissolve gold, then the king of metals. These metals are used chiefly for making jewelry, and for coinage.

The Strange Metals, Potassium and Sodium.—There are two strange substances which, though they do not appear to be metals, are metals just the same, and these are *potassium* (K) and *sodium* (Na).

They are the most active of any of the metals described above as you will see if you cut off a small piece of potassium and drop it into a dish of water. Hydrogen gas will be generated and the heat will be so great that the potassium will instantly catch on fire and burn with a series of *staccato* pops like a Lewis machine gun; at the same time it will be forced through the water in a zigzag line as shown in Fig. 40.

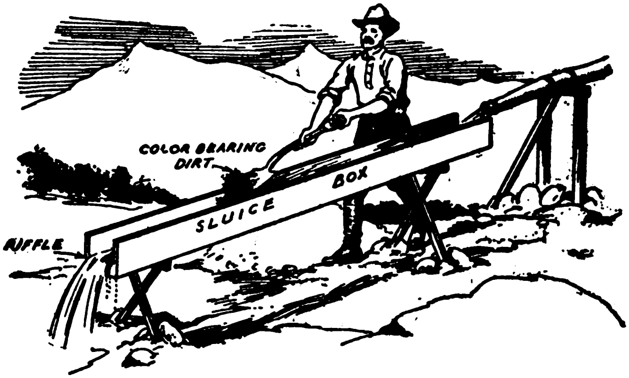


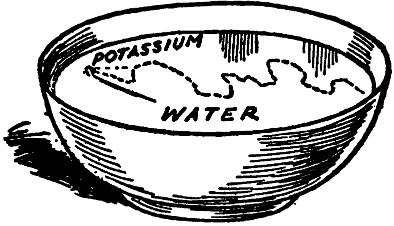
FIG. 39.—PLACER MINING FOR FREE GOLD.

Sodium does the same thing when dropped into water, but as the heat of the action is not so great the hydrogen does not catch on fire. The action of these two metals on water is as follows:

With potassium— $\overset{\text{Potassium}}{2K} + \overset{\text{Water}}{2H_2O} \rightarrow \overset{\text{Potassium hydroxide}}{2K(OH)} + \overset{\text{Hydrogen}}{H_2} \uparrow$

With sodium— $\overset{\text{Sodium}}{2Na} + \overset{\text{Water}}{2H_2O} \rightarrow \overset{\text{Sodium hydride}}{2Na(OH)} + \overset{\text{Hydrogen}}{H_2} \uparrow$

Thus you see that sodium and potassium are extremely active metals.



DOTTED LINE SHOWS PATH TAKEN BY POTASSIUM AS IT IS DRIVEN BY THE HYDROGEN GAS WHICH IS LIBERATED

FIG. 40.—HOW POTASSIUM ACTS ON WATER.

How to Tell a Metal.—Now you may ask, How can you tell which elements are metals, and which are non-metals? For instance, How are you to know that potassium and sodium are metals while sulphur and carbon are non-metals? Well, here is the answer:

When the oxides of metals are added to water they form the hydroxides or bases of the metals. Non-metallic elements are those whose oxides form acids when added to water.

Thus:

$$\overset{\text{Sodium oxide}}{Na_2O} + \overset{\text{Water}}{H_2O} \rightarrow \overset{\text{Sodium hydroxide}}{2Na(OH)} \text{ (Base)};$$

$$\overset{\text{Potassium oxide}}{K_2O} + \overset{\text{Water}}{H_2O} \rightarrow \overset{\text{Potassium hydroxide}}{2K(OH)} \text{ (Base)};$$

$$\overset{\text{Sulphur dioxide}}{So_2} + \overset{\text{Water}}{H_2O} \rightarrow \overset{\text{Sulphurous acid}}{H_2SO_3} \text{ (Acid)};$$

$$\overset{\text{Carbon dioxide}}{CO_2} + \overset{\text{Water}}{H_2O} \rightarrow \overset{\text{Carbonic acid}}{H_2CO_3} \text{ (Acid).}$$

The Activity of Metals.—Some metals have the property of being more *active* than others. Thus potassium acts directly with cold water to liberate hydrogen. Iron on the other hand will not act on cold water, but will liberate hydrogen from steam. Zinc will not liberate hydrogen from steam but will liberate it from acids while the precious metals will not liberate it at all.

The following list shows the order of activity of the metals, those at the top being the most active and those at the end of the list are the least active. It should be kept in mind, however, that the list shows the activity of the free metals only when they are combining with something else. Their activity depends on other things such as the solubility of the compounds for the free metal is not present at all.

ORDER OF ACTIVITY OF METALS

Potassium
Sodium
Magnesium
Aluminum
Manganese
Zinc
Chromium
Cadmium
Iron
Cobalt
Nickel
Tin
Lead
Hydrogen
Copper
Bismuth
Antimony
Mercury
Silver
Palladium
Platinum
Gold

Alloys of Metals.—There are many combinations that can be obtained by melting certain proportions of various metals together which produce mixtures whose properties are altogether different from either of the original metals which form it. Such mixtures of metals are known as *alloys* and they find as great, if not a greater, use in the commercial world than the free metals, of which they are formed.

Alloys of Iron and Steel.—*How Steel is Made.*——The *Bessemer process* of converting iron into steel consists of pouring the melted iron into an egg-shaped retort called a *converter*, see Fig. 41, and forcing a blast of air through it. The oxygen of the air unites with the carbon, silicon and other foreign substances in the iron and this leaves the latter nearly pure.

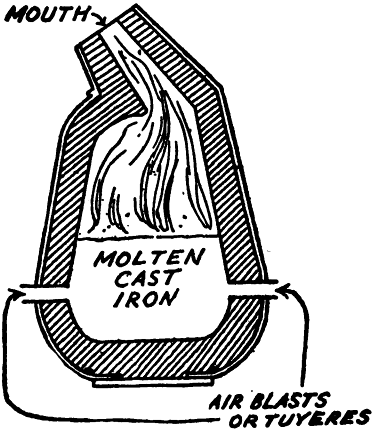


Fig. 41.—How Cast Iron is Changed into Steel by the Bessemer Converter.

The oxygen of the air burning in the melted iron raises the temperature to a point where the iron commences to *burn* and then cast iron containing carbon and sometimes other metals is put into the converter and when these are thoroughly mixed with the pure iron steel is formed.

Carbon Steel.—This has only a slight trace of *carbon* in it. When iron has a little carbon in it, it is hard and elastic, it can be welded and forged with ease and at the same time it can be *tempered*. It is called *carbon steel* and it makes an excellent material for tools and machinery of all kinds.

Manganese Steel.—However, there are alloys of steel that are even better than ordinary steel for special purposes. One of these is *manganese steel*, and this is made by melting iron containing 7 to 20 per cent of manganese with steel. It produces a steel of great hardness and it is largely used in the manufacture of steel jaws for rock crushers, for burglar-proof safes, etc.

Vanadium Steel.—Another kind of steel is *chrome vanadium steel*. This is a mixture of 1 per cent of chromium, and .15 per cent of vanadium with 98.85 per cent of steel. It has great *tensile* strength and can be bent double while cold without snapping. It is useful therefore for frames and axles of automobiles.

Tungsten Steel.—For *high speed* tools, that is tools which do not loose their temper even when red hot, *tungsten steel* is used, hence it is called *high-speed steel*. This contains 16 to 20 per cent of *tungsten*, .55 to .75 per cent of carbon, 2.5 to 5 per cent of chromium and .35 to 1.5 per cent of vanadium.

Nickel Steel.—*Nickel steel* contains 2 to 4 per cent of nickel, it resists the action of sea-water and is very hard and elastic. Hence it is used in making armorplate for battle-ships, for tanks, wire cables for underseas use and propeller shafts for ships.

Invar.—A steel known as *invar* contains 36 per cent of nickel and this alloy will not expand when it is heated; it is used for *meter sticks* for measuring, and *pendulum rods* for clocks which must always remain of exactly the same length.

ALLOYS OF COPPER.—As with steel, alloys of copper, have a wider use than the metal itself when free.

Bronze is one of the earliest alloys known. It contains

3 to 8 per cent of tin, 11 per cent of zinc, a little lead, and the rest copper. It is much harder than copper itself and it was used by primitive man during the *Bronze Age* for tools.

Brass is a more recent copper alloy. It contains 18 to 40 per cent of zinc and has a lower melting-point than copper. When hammered into thin sheets it is called *Dutch metal* and is similar to gold leaf which is used for gilding purposes, but, unlike gold, it tarnishes.

German silver, contrary to its name, contains no silver, but it is *German* all right, that is it is very much *à la camouflage*. It is made of copper and 19 to 44 per cent of zinc, and is used in the manufacture of imitation silver jewelry, for it looks much like silver.

Gun-metal contains 10 per cent of tin and 90 per cent of copper and has a beautiful brownish-black color. It is largely used for making art objects, medals, etc.

Amalgams.—Finally there is a special kind of alloys that are called *amalgams* and these are made by the union, or *amalgamation*, of mercury and some other metal.

All of the ordinary metals with the exception of iron and platinum will dissolve in mercury and this curious property of mercury for mixing with other metals is made use of in many ways in the arts, sciences and industries.

Amalgamating Battery Plates.—For instance, when the zinc plates of batteries are rubbed with mercury, or *amalgamated*, as it is called, the layer of mercury on the zinc prevents the acid of the battery from eating it away when it is not in use.

Dental Amalgam.—When powdered tin, silver or gold is made into an amalgam with mercury, the mass quickly hardens and hence, amalgams made with these metals are used by dentists for filling the cavities of your teeth, or someone's else teeth. Such amalgams expand and this makes them extremely suitable for this class of work.

Amalgam Silvering and Amalgam Gilding.—Silver and gold amalgams are used for silvering and gilding art objects. After the amalgam has been laid on, the mercury is driven off by the heat when a thin film of pure silver or gold is left on the surface. This process is called *amalgam silvering* or *amalgam gilding* according to the metal used.

CHAPTER VIII

PRACTICAL CHEMISTRY SIMPLY EXPLAINED

HOW DIFFERENT ELEMENTS COMBINE
WHAT VALENCE MEANS
WHAT ATOMIC WEIGHTS ARE
HOW TO CALCULATE WEIGHTS
ANOTHER PROBLEM IN WEIGHT
ABOUT BALANCING EQUATIONS
HOW TO CALCULATE VOLUMES
WHAT MOLECULAR WEIGHTS ARE

CHAPTER VIII

PRACTICAL CHEMISTRY SIMPLY EXPLAINED

In the chapters that have gone before I have told you many things that chemicals will do when they are combined in certain ways, but I did not tell you very much of the *how* or the *why* of these actions.

Unless you have some idea of how chemicals react on one another your chemistry will not be of any great value to you, that is if you intend to become a real chemist, for the chemistry of to-day has all to do with the theory of things.

For instance a practical chemist must not only know that when hydrochloric acid acts on zinc, hydrogen is given off, but he must be able to calculate just *how much* hydrogen will be given off by a certain amount of these substances, and the reverse, that is how much hydrochloric acid and zinc he will need to obtain a given amount of hydrogen.

There are hundreds of other little calculations such as these that he is constantly called upon to make and if you learn how to make them now, and you can by *studying* this chapter, you will be well advanced along the road to the successful practice of chemistry.

How Different Elements Combine.—The first thing that you have to learn is why certain elements combine in a certain way. Thus why do zinc and chlorine combine with one part of zinc and two of chlorine to form zinc chloride, $ZnCl_2$, instead of one part of zinc to some other number of parts of chlorine, or the other way about?

To find out why, consider the action of three metals on hydrochloric acid, namely sodium, zinc and aluminum. As you already know, all of these metals will liberate hydrogen from an acid and form a salt of the metal thus:

$$Na + HCl \rightarrow NaCl + H\uparrow;$$

$$Zn + 2HCl \rightarrow ZnCl_2 + 2H\uparrow;$$

$$Al + 3HCl \rightarrow AlCl_3 + 3H\uparrow.$$

From this you will see that sodium liberates 1 part of hydrogen from the acid, zinc liberates 2 parts and aluminum liberates 3 parts. Also you will notice that the sodium is able to combine with 1 part of chlorine, zinc with 2 parts of chlorine, and aluminum with 3 parts of chlorine.

For this reason hydrogen is taken as the *unit of combining power*, that is to say anything that can take the place of 1 part of hydrogen, or liberate 1 part of hydrogen as the sodium does is said to have a combining power of 1. Since zinc takes the place of 2 parts of hydrogen it is said to have a combining power of 2, while aluminum has a combining power of 3.

What Valence Means.—This combining power of an element is called its *valence*. Thus sodium, hydrogen, and chlorine have a valence of 1. Zinc, calcium and oxygen have a valence of 2, aluminum has a valence of 3, while carbon has a valence of 4. I might add here that (OH) the hydroxile radical has a valence of 1.[1]

Now knowing what the valences or combining powers of two elements are you can easily tell how they will combine, that is what the formula will be, and what compound they will make.

Thus if you want to know how H and Cl will combine or what the formula is all you have to do is to remember

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[1] For the valences of the different elements see Appendix IV.

that H has a valence of 1 and Cl has a valence of 1. Consequently 1 part of H can combine with 1 part of Cl, which makes HCl or *hydrogen chloride.*

So, too, if you want to know how Zn and O will combine just bear in mind that both elements have a valence of 2, and you know that 2 parts of zinc can hold 2 parts of oxygen. Since this is true you also know that 1 part of zinc can hold 1 part of oxygen and hence the formula is ZnO or *zinc oxide.* See?

Suppose, though, that you want to know how Zn and Cl combine. Now Zn has a valence of 2 and Cl has a valence of only 1. Consequently if you are going to have a compound you must take 2 parts of chlorine in order to combine with the zinc which has the power to hold 2 parts of any element with a valence of 1. The formula will then be $ZnCl_2$ or *zinc chloride.*

Likewise since aluminum has a valence of 3 and chlorine of 1, it will take 3 parts of chlorine to combine with 1 of aluminum, for the aluminum can hold 3 parts of chlorine, thus $AlCl_3$ or *aluminum chloride.*

Now, however, we come to one a little harder. Suppose we want to know how aluminum and oxygen will combine, aluminum having a valence of 3 and oxygen of 2. In order to make the formula of a compound where the valences of the elements are unequal the rule is this:

Write down the first element (aluminum in this case) with the valence of the second element after it (2) and then write down the second element (oxygen) with the valence of the first element (3) after it thus, Al_2O_3, and this is the formula for *aluminum oxide.*

You will see that there are 2 parts of aluminum and you know that each part has a valence of 3 or the 2 parts of aluminum have a total valence of 2×3 or 6. Also you will see that there are 3 parts of oxygen and that each part has a valence of 2; the total valence of the oxygen then is 3×2 or 6.

From this you observe that both elements have the same *total* valence and this is true in every compound containing two elements. If you will practise combining elements according to their valences you will find that you have no trouble in writing the formula for any compound you want.

What Atomic Weights Are.—After you have learned to write the formula of any compound the next thing to do is to learn to calculate the weights of the materials needed in chemical actions.

In *Appendix* B opposite each of the symbols of the element you will see a number. Thus opposite the element *chlorine* you will find written the number 35, and opposite the element *hydrogen* the number 1.

Now this simply means that when hydrogen and chlorine combine, 1 gram of hydrogen will combine with 35 grams of chlorine. These numbers are called the *atomic weights* of the elements. Thus the formula for *hydrogen chlorine* you know from the valences of its parts to be HCl and since 1 gram of hydrogen combines with 35 grams of chlorine the weight of this compound is 35+1 or 36 grams. This weight is known as the *formula weight.*

How to Calculate Weights.—Suppose now that you are going to add 100 grams of *sodium* to *hydrochloric acid* and you want to know how much *hydrogen* the *sodium* will set free from the acid.

In making calculations about weights there are three things to do first and these are (1) to write the formula for the action, (2) to write the atomic weights under the substances you are going to work with and (3) to write the real weights you are using over the substances you are going to work with.

When you are trying to find the *real* weight of a substance write an x over that substance thus:

$$\underset{23}{\overset{100\text{ grams}}{\mathrm{Na}}}+\mathrm{HCl}\rightarrow\mathrm{NaCl}+\underset{1\text{ (atomic weight)}}{\overset{\text{(real weight) }x}{\mathrm{H}}}$$

You will observe that you do not need to bother with the HCl and the NaCl because you are not interested in them, also that you write an x over the H because that is the substance whose weight you want to find.[1]

Next write down the atomic weights and the real weights as follows:

$$23 : 1 = 100 : x.$$

This is the same as reading the problem, 23 grams of sodium give 1 gram of hydrogen therefore 100 grams of sodium will give x or how many grams of hydrogen? This is known as a *proportion* and in order to find what x is equal to use the following rule:

Multiply the two end members (23 and x) together (these are called the *extremes*) and set the product down thus

$$23 \text{ times } x = 23x;$$

then multiply the two middle numbers (1 and 100) together (these are called the *means*) and set this product down equal to the product of the extremes thus

$$23x = 100.$$

Now in order to find x you have only to *divide* 100 *by* 23 when the answer or *quotient* will be thus:

$$100 \div 23 = 4.34 \text{ grams of hydrogen.}$$

Where 4.34 grams of hydrogen is x and is the amount that will be set free by 100 grams of sodium acting on hydrochloric acid.

Another Problem in Weight.—Just for the sake of practise take another problem of the same kind. Suppose you want to know how many grams of hydrochloric acid are necessary to make 66 grams of aluminum chloride.

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[1] An unknown quantity is always represented by x.

First write the formula with their real and atomic weights.

$$\overset{x\text{ grams}}{Al+3HCl} \rightarrow \overset{66\text{ grams}}{AlCl_3+3H}.$$

Now HCl is made up of 1 gram of hydrogen (atomic weight) and 35 grams of chlorine (atomic weight) or HCl weighs 35+1 or 36 grams. But you have 3HCl or 3×36 or 108 grams, and this equals the total weight. Also aluminum chloride is made up of 27 grams (atomic weight) of aluminum and 35 grams of chlorine (atomic weight) but since you have Cl_3 the total weight of chlorine is 3×35 or 105 grams and the total weight of aluminum chloride is 27+105 or 132 grams.

So writing these weights under the compounds you have

$$\underset{108\text{ grams (at. wt.)}}{\overset{x\text{ grams (real weight)}}{Al+3HCl}} \rightarrow \underset{132\text{ grams (at. wt.)}}{\overset{66\text{ grams (real weight)}}{AlCl_3+3H}}$$

you will see that in this case you do not have to bother with the Al or the H because you are not interested in them.

Next write down the proportion by writing the atomic weights first and then the real weights and then you will have:

$$108 : 132 = x : 66.$$

Now the problem reads, *if* 108 *grams of HCl give* 132 *grams of $AlCl_3$ how many* (x) *grams of HCl will be needed to give* 66 *grams of $AlCl_3$.*

Multiplying the means you have

$$132 \text{ times } x = 132x$$

Multiplying the extremes you have

$$108 \text{ times } 66 = 7128$$

Then

$$132x = 7128$$

or

$$x = 7128 \div 132 = 54 \text{ grams of hydrochloric acid.}$$

That is 54 grams of hydrochloric acid will be needed to make 66 grams of aluminum chloride.

Thus, you see, that with a little thought and the use of the methods I have shown you, you can calculate the weights of materials you need with very little trouble.

About Balancing Equations.—Finally a few words more on *making equations*. You may wonder why in the formula which follows:

$$Al + 3HCl \rightarrow AlCl_3 + 3H$$

you need to use 3HCl instead of only 1HCl as in the formula,

$$Na + HCl \rightarrow NaCl + H.$$

The reason is this: You know from your experiments that when you add hydrochloric acid to aluminum you get aluminum chloride and hydrogen. Also recalling your valences you know that the formula for aluminum chloride must be $AlCl_3$. So the first thing for you to do is to write the equation thus:

$$Al + HCl \rightarrow AlCl_3 + H.$$

From the above you instantly see that you have 3 chlorines on the *right side* of the equation while on the *left* you have only 1. Since you will need 3HCl in order to furnish the 3 chlorines for the $AlCl_3$ you must now write the equation

$$Al + 3HCl \rightarrow AlCl_3 + H.$$

But now you observe that you have 3 hydrogens on the *left* and only 1 on the *right*. So you try again and write the equation like this

$$Al + 3HCl \rightarrow AlCl_3 + 3H$$

and you will find that you will have the same valences on *both sides* of the equation and this means that it is now correct.

This process is called *balancing the equation* and you should practise writing and balancing equations for this is an important part of a chemist's business. After some practise you will get it down so that you can balance an equation in your head instead of writing it out three or four times on paper as I have here to show you how it is done.

How to Calculate Volumes.—Having found out how to calculate the weights of the chemicals needed for and produced by chemical actions you should learn how to calculate the *volumes*, for in chemistry it is often necessary to know not only what weight a product has but also how much *space*, that is volume, it takes up.

While the making of equations as explained above is all right for calculating weights, the method cannot be used for calculations of volume. The thing you want to keep in mind is that the only time it is needed and possible to calculate volumes is when one or more of the substances used or produced is in a gaseous state.

Thus in the following formula

$$2Al+3Cl_2 \rightarrow 2AlCl_3$$

it is not possible to calculate the volume of aluminum used, or of the aluminum chloride formed, for both of these substances are solids. The *volume* of the chlorine gas used, however, can be calculated for it is in gaseous form. Now let us see why this is possible.

What Molecular Weights Are.—In the first place the above equation is what is known as a *molecular equation*, that is, it takes into account the molecules of the gas as well as the atoms.

Now when you calculated weights you learned how to write *atomic equations* because the calculations of weight depend on the *atomic weights*, but the calculation of volume depends on the *molecular weights* of substances.

Thus the atomic equation for the combination of aluminum and chlorine is

$$Al+3Cl \rightarrow AlCl_3.$$

But 3Cl represents the number of atomic weights of chlorine and not the molecular weight.

The molecular weight of a gas is the weight of that gas needed to fill a cube of 22.4 *liters*[1] capacity. For all gases this weight is equal to the atomic weight times the number of atoms in a molecule of the gas.

Such gases as oxygen, chlorine and hydrogen contain 2 atoms in every molecule hence their molecular formula is not O, Cl and H, but O_2, Cl_2 and H_2, where the two represents the number of atoms in the molecule.

But since you are calculating volume and since volume depends on the molecular formula you must change the *atomic* equation to a *molecular* one. Chlorine is the gas you want to find the volume of, so in writing the formula you must first write your chlorine in molecular form and then the valance in the equation after, in this fashion:

$$Al+3Cl_2 \rightarrow AlCl_3.$$

The $3Cl_2$ now represents the volume of chlorine needed in this equation. However, you now have 3×2 or 6 chlorines on the left and only 3 on the right. Hence it is necessary to multiply the right hand side of the equation by 2 in order to give 6 chlorines and so make it balance the left, thus:

$$Al+3Cl_2 \rightarrow 2AlCl_3.$$

But you will see now that there are 2 aluminums on the right and only 1 on the left; now to offset this multiply the aluminum on the left by 2 when the equation

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[1] See Chapter III.

will be the molecular one needed for calculating volumes and will be balanced thus:

$$2Al+3Cl_2 \rightarrow 2AlCl_3.$$

You are ready now to calculate volumes. Suppose you want to find out what volume of chlorine is needed to give 50 grams of aluminum chloride. The first thing to do is to write the equation with the molecular weights and volumes under it, like this:

$$\begin{array}{ccc} 2Al+ & 3Cl_2 & \rightarrow 2AlCl_3 \\ & 3x22.4 \text{ liters} & 2x133.48 \text{ grams} \\ & x \text{ liters} & 50 \text{ grams} \end{array}$$

Where 3x22.4 liters is the formula volume of chlorine, because, as I told you, Cl_2 represents the molecular weight, that is the weight that can fill 22.4 liters, and since you have $3Cl_2$ the volume will be 3x22.4 liters:

x liters is what we want to find, that is, how many liters of chlorine will be needed to make 50 grams of aluminum chloride:

2x133.48 is the molecular weight of the aluminum chloride (the atomic weight is as follows: $Al=27.1$, $Cl_3=3x35.46=96.38$ and $27.1+96.38=133.48$). Since there are 2 volumes of $AlCl_3$ we have 2x133.48 grams.

And, finally, 50 grams is the weight of aluminum chloride you want to make. If now we make a proportion between the formula weights and volumes and the actual weights and volumes you will be able to calculate the volume of chlorine needed thus:

$$2x133.48 : 3x22.4 = 50 : x \text{ liters},$$

or

$$267.96 : \quad 67.2 = 50 : x \text{ liters};$$

multiplying the extremes we have

$$267.96x,$$

multiplying the means we have

$$3360.$$

Setting them equal you have

$$267.96x = 3360,$$

or $$x = 3360 \div 267.96 = 12.54 \text{ liters},$$

and 12.54 liters of chlorine is the volume needed to combine with aluminum to form 50 grams of aluminum chloride.

CHAPTER IX

SOME BRILLIANT CHEMICAL TESTS

TESTING WITH INDICATORS
    (*a*) Litmus paper
    (*b*) Phenolphthalein
    (*c*) Methyl Orange, and
    (*d*) Congo Red
HOW TO ANALYZE SUBSTANCES
    The Wet Method of Analysis
        Steps in Making Analysis
A TEST FOR IODINE
TESTS FOR PHOSPHORIC ACID
HOW TO MAKE PURPLE OF CASSIUS
A TEST FOR DICHROMIC ACID
THE TEST FOR SULPHUR
A TEST FOR MERCURIC IODIDE
HOW TO MAKE SYMPATHETIC INK

CHAPTER IX

SOME BRILLIANT CHEMICAL TESTS

VERY often in chemistry you will have a substance or a compound whose nature you do not know but which you would like to know; or you may have a compound that you suspect has some ingredient in it which you do not know but would like to know.

To find out merely whether a substance or a compound is an acid or an alkaline is a very simple matter but to ascertain whether a compound contains say *nickel sulphide* or the salts of some other metal is not so easy though you can do it if you know how and it is my idea to let you into the secret of it.

Testing with Indicators.—*Litmus Paper.* In the chapter on *Acids, Bases* and *Salt* you will remember I mentioned the use of *litmus paper* for testing acids and salts and that this test paper *indicates* whether the solution is an acid or an alkaline by the color it takes on.

Litmus is a blue dyestuff made by fermenting *lichens*, and it was first obtained by Boyle. *Litmus paper* is made by dipping absorbent paper into a solution of litmus and then drying it. When it is dipped into an acid solution it turns red while it remains blue when dipped into an alkaline solution. Hence, it is called an *indicator*.

PHENOLPHTHALEIN. — (Pronounced *fen-ol'-tha-lean*). — This is another indicator for acids and salts. It is a colorless substance and, different from litmus, it turns red when it comes in contact with alkaline solutions but it remains colorless in the presence of acids.

Phenolphthalein is also used as an indicator of the poles of an electric circuit, that is it shows which is the positive and negative wires. A *polarity indicator* consists of a small glass tube filled with a solution of phenolphthalein and in each end of which a bit of metal is sealed as shown in Fig. 42. When the indicator is connected across the circuit the solution will turn red around the negative end of the wire.

METHYL ORANGE and CONGO RED.—When *methyl orange* is added to acids it changes to a red color and when added to alkalines it gives a yellow color. When *Congo red* is added to acids it turns a deep blue while with alkalines it remains red.

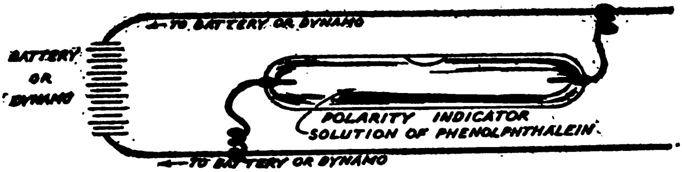

FIG. 42.—A POLARITY INDICATOR, WHICH INDICATES INSTANTLY THE NEGATIVE AND POSITIVE POLES OF A BATTERY.

How to Analyze Substances.—Now we'll do something a little harder and that is find out if a substance or a compound contains a certain unknown substance or compound. To do this you test them with *reagents*, that is with other chemicals which will act on them and at the same time you will get some very pretty color effects.

Now the reduction of a chemical compound into its original elements, so that one or more of them can be recognized, is called *qualitative analysis*, because you find out the *quality* of it, that is what substances it contains. But to find out just how much of one or more of the elements there is in a compound, that is in what proportions they have combined, is called *quantitative analysis*, and this is even harder.

In the first part of this chapter I shall explain what is known as the *wet method* of making *qualitative* analysis. By this method you can tell what the *positive ion* is, which means that part of the salt that carries the *positive charge* of electricity. Since the positive ion of a salt is always a metal you can take a compound and tell what metal it is a salt of.

THE WET METHOD OF ANALYSIS.—This method of *qualitative analysis* for metals, that is finding out what the metal is in a compound, is made by dissolving the metallic salt in water and then adding certain chemicals which form *precipitates* of various colors and these fall to the bottom of the test tube or beaker.

Stated more fully analyses[1] of this kind are made by adding *sulphides* to the above solutions, since the sulphides of the metal will then be precipitated and these always have colors by which you can easily know them.

There are four steps in making a test of this kind (see Fig. 43), and these are often necessary for the salts of the metal may be of such a nature that the metal will not be recognized until all four of them have been applied and more often the salt which you are examining will not be a pure specimen but will be mixed with other salts. Before you take the first step dissolve the salt thoroughly in water.

The First Step.—First add hydrochloric acid to the solution you are going to test. If silver, mercury or lead are present the *insoluble chlorides* of these metals, namely, $AgCl$, $HgCl$ and $PbCl_2$, will be formed and thrown down, that is precipitated. All of these precipitates are white and to tell them apart you must make a further test.

Filter off the liquid and to the precipitate add *ammonium*

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[1] This is the plural of analysis.

hydroxide. If the precipitate is *silver chloride* it will turn *black,* while if it is *lead chloride* it will *not dissolve* or change color at all.

The Second Step.—Should, however, no precipitate be formed at all in the first step, and if this is the case, generate some *hydrogen sulphide* as described before and pass it through the solution which the hydrochloric acid you added to it has made acid, that is, the hydrochloric acid has given acid properties to it.

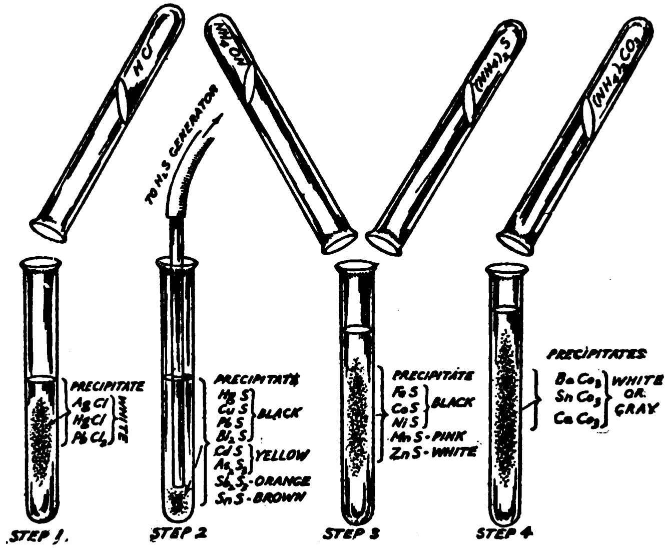

FIG. 43.—THE WET METHOD OF ANALYSIS.

The sulphides of those metals which are insoluble in active acids will now be precipitated if any of these metals are present. These sulphides are *mercuric sulphide* HgS, *cupric sulphide* CuS, *lead sulphide* PbS, *bismuth trisulphide* Bi_2S_3, *cadmium sulphide* CdS, *arsenious sulphide* As_2S_3, *antimony trisulphide* Sb_2S_3 and *tin sulphide*, or *stannous sulphide*, SnS, as it is called.

The first four of these sulphides are *black* or brown, the next two are *yellow* and the last two are *orange* and *brown*, respectively. Now these precipitates can be told one from the other like this:

The last three named pass into solution, that is they dissolve when warmed with yellow *ammonium sulphide*, while the first five sulphides named will remain unaffected by this treatment. However, with the exception of mercuric sulphide, all of the first five sulphides named will react when mixed with *hot* nitric acid.

The Third Step.—If now these precipitates, should they be formed, are filtered out of the solution, or if they are not formed and ammonium hydroxide is added to neutralize the acid, when the ammonium sulphide is added the sulphides which are insoluble in water will be precipitated.

These precipitates are *iron sulphide* FeS, *cobalt sulphide* CoS, *nickel sulphide* NS, *magnesium sulphide* MnS and *zinc sulphide* ZnS. Of these precipitates the first three are *black*, the next is *pink* while the last one is *white*. And now you are ready for the final step. If any one of these precipitates has been formed, filter it out, or if not go on to step four.

The Fourth Step.—Add *ammonium carbonate* to the solution when metals whose carbonates are insoluble will be precipitated out. These are *barium carbonate*, *strontium carbonate* and *calcium carbonate*.

There will be left now only the *salts of sodium*, *potassium* and *ammonium* in the solution. Hence, you see that you will be able from these precipitates to recognize nearly all of the heavier metals. These various steps are shown in Fig. 43.

Now while I have picked out nearly all of these tests, because of their striking color effects, you can see that if used after the fashion described they will prove serviceable as well. There are some other tests that I am going to tell you about which are quite interesting and useful but like those that have gone before I have chosen them because they produce brilliant colors.

The Test for Iodine or Starch.—The first of these is the *test for iodine* or for *starch.* Drop a few *crystals of sublimed iodine* into water and shake them well. After a little time some of them will dissolve and the solution will take on a faintly perceptible brown color.

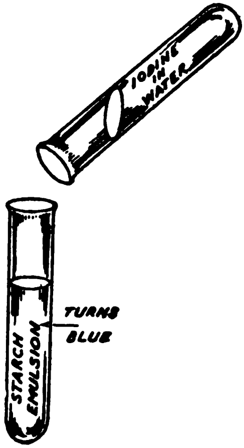

FIG. 44.—A TEST FOR IODINE OR STARCH

Now make a *starch emulsion.* To do this heat a beaker of water to the boiling-point and then take it from the flame. Stir in starch until the liquid becomes opaque, that is so that you can not see through it, and while still hot filter it saving the clear *filtrate,* that is the part which goes through the filter paper.

When the filtrate has cooled it should look like ordinary water. If it is cloudy repeat the operation described above until you get a filtrate that is clear. If now one drop of the slightly colored iodine water is added to the filtered starch emulsion and shaken well, the whole solution will turn a beautiful deep blue. (See Fig. 44).

This blue is not the result of some chemical compound being formed as you might have supposed, but is merely the result of the iodine being absorbed by the starch the minute particles of which are held in *colloidal*[1]

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[1] When a finely divided substance is suspended in a solution it is called a *colloid.*

suspension in the solution. The above test is equally good for iodine or starch should the presence of either be suspected and the blue is a characteristic color of them which you will always be able to recognize.

Often when a compound is suspected of containing iodine a solution is made of it and *chlorine gas* is passed through it, when, if there is any iodine present the chlorine will set it free from the compound. The starch emulsion test can then be applied and it can be discovered whether or not the compound contained iodine. *Bromine* is another element known as a *halogen*[1] which is very like unto and is of the same family as chlorine; it will also liberate iodine from its compounds when the latter are in solution.

Tests for Phosphoric Acid.—It is also good to know how to make tests for those acids which the element phosphorus forms. Now phosphorus forms three chief acids and these are *orthophosphoric acid*, *pyrophosphoric acid* and *metaphosphoric acid*. Since there are three of these acids two tests are necessary in order to tell them one from the other.

To a solution which you suspect may have one of the three acids in it, add *silver nitrate* first. If the solution contains orthophosphoric acid a *yellow* precipitate of *silver orthophosphate* will be thrown down, or precipitated.

If however either pyrophosphoric or metaphosphoric acids are present then a *white* precipitate of silver *pyrophosphate* or *silver metaphosphate* will be formed, and hence it is necessary to distinguish between these latter two.

If to the solution of acid a clear solution of albumen, such as the white of egg, is added it will be *coagulated*,

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[1] The word *halogen* means *salt former* and as fluorine, bromine, chlorine, and iodine directly unite with a large number of metallic elements, they are called the *halogens*.

that is it will look curdled if metaphosphoric acid is present whereas pyrophosphoric acid will not affect it, all of which is shown in Fig. 45.

This is the way then that you can distinguish the *pyro* from the *metaphosphoric acid.*

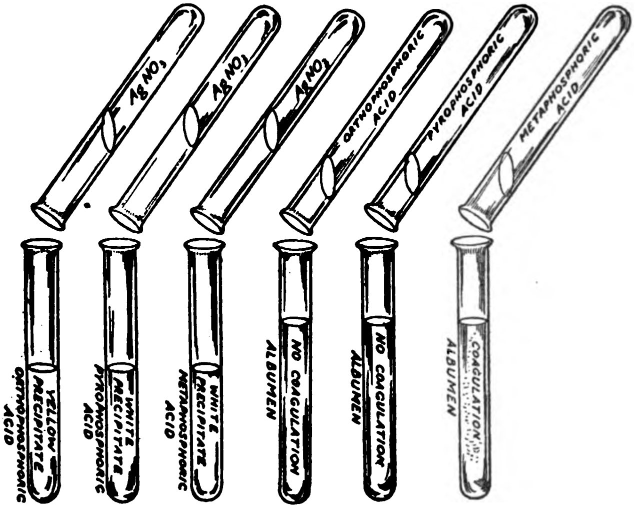

FIG. 45.—THE PHOSPHORIC ACID TEST

Orthoarsenic acid is an acid similar to orthophosphoric acid and when silver nitrate is added to it it gives a characteristic *chocolate brown* precipitate by which you can always know it.

How to Make Purple of Cassius.—One of the prettiest tests you can carry out is to make what is known as the *purple of Cassius.* To a solution of *sodium chloraurate* add a little *stannous chloride* solution. Sodium chloraurate is a compound composed of sodium, chlorine and gold. The stannous chloride is a very active reducing agent and consequently reduces the sodium chloraurate setting the gold free and forming sodium chloride and *stannic chloride,* thus

$$\overset{\text{Sodium chloraurate}}{2NaAuCl_4} + \overset{\text{Stannous chloride}}{3SnCl_2} \rightarrow \overset{\text{Sodium chloride}}{2NaCl} + \overset{\text{Stannic chloride}}{3SnCl_4} + \overset{\text{Gold}}{2Au\downarrow}$$

Now the stannic chloride is *hydrolized* by the water, that is the water acts upon it and the stannic acid is precipitated. The gold which was set free was in such minute particles that it did not sink to the bottom of the beaker as does an ordinary precipitate. Instead the particles remained suspended throughout the liquid and they are so fine that filtering will not remove them either. A fine suspension of this kind in a liquid is called a *colloidal suspension.*

Here then we have a colloidal suspension of pure gold and a precipitate of stannic acid. The suspension of gold *disperses,* that is spreads itself on the precipitate of stannic acid which is of a colloidal nature itself and it is the light passing through this curious combination that gives the beautiful deep color known as the purple of Cassius.

A Test for Dichromic Acid.—One of the prettiest blues you ever saw is formed when *a dichromic acid* test is under way, the formula for which is $H_2Cr_2O_7$. When a trace of a *dichromate* is suspected of being present in a solution it can be detected as follows:

Add a drop of the solution to a solution of *hydrogen peroxide* to which enough *sulphuric acid* has been added to make the hydrogen peroxide slightly acid, or to turn blue litmus paper red. The acid reacts on the hydrogen peroxide and gives free dichromic acid if it is present, thus:

$$\overset{\text{Sulphuric acid}}{H_2SO_4} + \overset{\text{Potassium dichromate}}{K_2Cr_2O_7} \rightarrow \overset{\text{Dichromic acid}}{H_2Cr_9O_7} + \overset{\text{Potassium sulphate}}{K_2SO_4}$$

At the same time a *perchromic acid* is formed by the action of the hydrogen peroxide upon the free dichromic acid. This substance gives the beautiful *blue* color that makes this test a distinctive one and which is used by all chemists.

The Test for Sulphur.—Another very interesting test is that for *sulphur*. Although sulphur is yellow in color still this is not enough to know that a substance which looks like it is sulphur. One of the easiest and quickest ways of finding out whether a substance is sulphur or not is to place a little in a dry test tube and heat it.

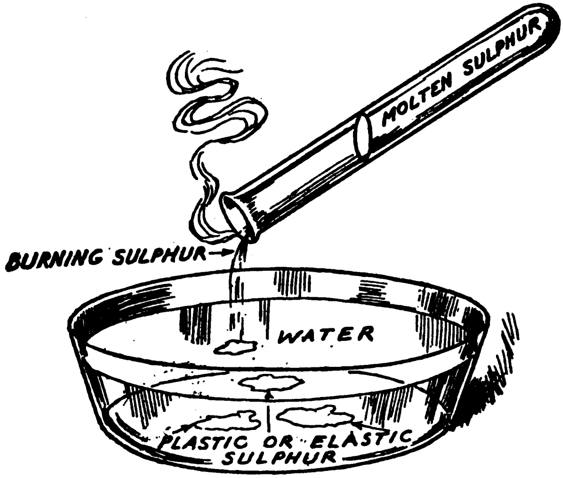

Fig. 46.—How Plastic Sulphur is Made

If it is sulphur it will shortly begin to melt turning a pale *straw color*. Keep on heating it when it will turn a dark brown and soon get solid. Finally this solid mass will melt and a thick reddish brown fluid will be formed.

Now ignite the fluid by holding the mouth of the test tube containing it across the flame for a second.

The sulphur will catch on fire, and then pour the burning solution into cold water. Upon striking the water it sinks slowly to the bottom and when you fish it out you will find that it has turned into a rubber-like mass which is elastic and is known as *plastic sulphur*, as shown in Fig. 46. After letting this mass stand for several days it will form crystals. Of course this test is only good where pure uncombined sulphur is to be tested.

A Test for Mercuric Iodide.—Another curious test is that for *mercuric iodide*, the formula for which is HgI_2. This is a red powder formed by the addition of *potassium iodide* to *salts of mercury*.

Put a little in a test tube and heat it as you did the sulphur. Very soon it will turn yellow and then it will melt. Have a glass rod handy, and pour the melted mass out of the test tube onto a piece of paper. After cooling for a fraction of a minute it will form yellow crystals.

Now touch the yellow crystals with the glass rod when they will turn red or a *scarlet* color similar to that of the powder before it was melted.

How to Make Sympathetic Ink.—Here is a little experiment and though it is not a test it is both curious and amusing. This is the making of invisible ink, that is ink which can only be seen when you want it to be seen and hence it should be included in every *Lover's Complete Letter Writer*.

When *cobaltous oxide* is dissolved in hydrochloric acid red prisms, or crystals of *cobaltous chloride* are formed, thus:

$$\underset{\text{Cobaltous oxide}}{CoO} + \underset{\text{Hydrochloric acid}}{2HCl} \rightarrow \underset{\text{Cobaltous chloride}}{CoCl_2} \downarrow + \underset{\text{Water}}{H_2O}.$$

Now if these crystals are dissolved in water and the solution is used as an ink the writing will become invisible as soon as the ink dries. When you want to read it all you have to do is to warm the writing, when it becomes blue and hence visible because you have driven the water out of it. Shortly after it will again absorb water from the air and become invisible. The heating process can then be carried out again. This kind of ink is known as *sympathetic ink*.

CHAPTER X

CHEMISTRY IN AND AROUND THE HOUSE

How Hard Water Wastes Soap and Makes Work
How to Make Hard Water Soft
On the Use of Washing Soda
The Value of Household, or Aqua Ammonia
How Borax Softens Water
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About Bleaching and Bleaching Compounds
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    Red Lead; White Lead
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CHAPTER X

CHEMISTRY IN AND AROUND THE HOUSE

CHEMISTRY plays a very considerable part in and around the house and yet few people pay any attention to it. A working knowledge of a few simple chemicals is quite indispensable in any well-regulated household for it not only saves much time and hard work but it helps to cut down the HCL, which in this case is not a chemical formula but is the short way of saying the *high cost of living*, and this is well worth taking into account.

How Hard Water Wastes Soap and Makes Work.—Since cleanliness is next to godliness, the first thing you should know about is *hard water*, its action on soap and how you can make it soft.

There are two kinds of hard water and these are (1) *temporary* hard water, and (2) *permanent* hard water. Temporary hard water is caused by ordinary water flowing over limestone from which it takes up *calcium carbonate*, which is *limestone*, while permanent hard water contains *calcium sulphate*, that is *gypsum*.

To show how hard water of either kind wastes soap and makes work let me point out that 6 pounds of soap are wasted for every 100 gallons of water used because it takes that much to precipitate the calcium salts before the soap can actually begin to do its work of cleaning.

More than this even after the soap has softened the water the calcium salts are acted on by the acids in the soap and form a precipitate which is very sticky and which is likely to soil the things that are being washed.

How to Make Hard Water Soft.—It stands to reason then that the thing to do when you intend to wash clothes in hard water is to get rid of the hardness before you try to use soap in it.

You can easily tell whether or not water is hard by finding out how easy or hard it is to make a lather of it with good soap and whether or not the lather is lasting. If the soap will not lather freely and if it does not last the water you have used with it is hard.

Slaked lime is often used to soften hard water on a large scale for factories but for household use there are other substances, which are cheaper and quicker in the long run and these are (1) *washing soda*, (2) *household* or *aqua ammonia* and (3) *borax*.

On the Use of Washing Soda.—*Soda, sal soda* or *washing soda* as it is variously called, is simply *sodium carbonate* (Na_2CO_3). Now when a small quantity of washing soda is added to hard water, whether its hardness is temporary or permanent, or both, it removes it like this:

$$\overset{\substack{\text{Washing}\\ \text{soda}}}{Na_2CO_3} + \overset{\text{Hardness}}{Ca(HCO_3)} \rightarrow \overset{\substack{\text{Calcium}\\ \text{carbonate}}}{CaCO_3} \downarrow \overset{\substack{\text{Sodium}\\ \text{bicarbonate}}}{2NaHCO_2}.$$

Both the calcium carbonate and the sodium carbonate are insoluble, that is they will not dissolve, and are precipitated to the bottom of the tub, or other vessel holding the water. When this action takes place the soap has a chance to do its full duty and hence the family washing is not only made a little cheaper but a great deal easier.

The Value of Household or Aqua Ammonia.—*Household* or *aqua ammonia* is made by heating *ammonium chloride* which breaks down and gives off *ammonia gas*. The ammonia gas is then passed into water in which much of it dissolves, when it reacts with the water to form *ammonium hydroxide* (NH_4OH), a base, and it is to this substance that ammonia owes its cleaning power in cutting grease and dirt.

When NH_4OH is added to hard water it removes both the temporary and the permanent hardness, thus:

$$\overset{\substack{\text{Ammonium}\\ \text{hydroxide}}}{2NH_4OH} + \overset{\text{Hardness}}{Ca(HCO_3)_2} \rightarrow \overset{\substack{\text{Calcium}\\ \text{carbonate}}}{CaCO_3\downarrow} + \overset{\substack{\text{Ammonium}\\ \text{carbonate}}}{(NH_4)_2CO_3} + \overset{\text{Water}}{2H_2O}$$

The ammonia does not act on the calcium sulphate, or permanent hardness, but the ammonium carbonate which is formed when it acts on the temporary hardness does and as a result two more precipitates are formed and the water is freed of its hardening substance, thus:

$$\overset{\substack{\text{Ammonium}\\ \text{carbonate}}}{(NH_4)_2CO_3} + \overset{\text{Hardness}}{CaSO_4} \rightarrow \overset{\substack{\text{Calcium}\\ \text{carbonate}}}{CaCO_3} + \overset{\substack{\text{Ammonium}\\ \text{sulphate}}}{(NH_4)_2SO_4\downarrow}$$

How Borax Softens Water.—The formula for *borax* is $Na_2B_4O_7$ and when it is added to hard water it reacts with the water just as ammonia *gas* does and gives *sodium hydroxide.* The sodium hydroxide then acts on the hardness exactly as does the ammonium hydroxide just described.

It is a common mistake to suppose that soda and borax have any bleaching or whitening action on goods. What these two substances really do is to precipitate any salts of iron that may be present in the water and thus prevent them from acting with the acids of the soap when they would stain the clothes a reddish brown.

Soda, ammonia and borax will remove animal or vegetable oils and grease but have no effect whatever on mineral oils.

How Soaps are Made.—Having found out something about the cleaning powers of soap it will be a good scheme to learn how it is made. In making soap, *lard* or *beef fat* is first mixed with hot sodium *hydroxide* solution and heated with alcohol.

The fat then spreads through the solution in little drops and what is known as an emulsion is formed;

the alcohol, which is the hydroxide, acts on the fat and it decomposes it into *sodium palmitate, sodium stearate, sodium oleate,* and *glycerine.*

Common table salt is next added to the mixture when these *sodium* salts *coagulate,* that is they get like jelly, and separate into a floating layer on top of the liquid, while the glycerine is contained in the salt solution.

This coagulating process by means of table salt (see Fig. 47), is known as *salting out* and it is also used in many other processes besides that of soap making. The coagulation is then skimmed off of the surface of the liquid and when cold it gets hard and this is the *soap.*

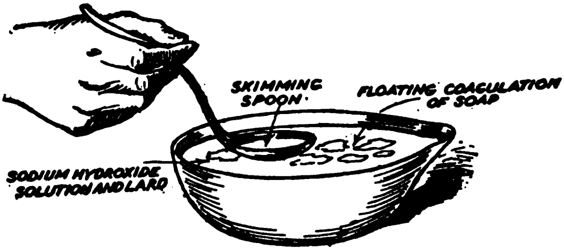

Fig. 47.—Coagulation of Soap after "Salting out."

The Kinds of Soap we Use.—When you want to make that kind of soap which is called *soft soap,* heat *potassium hydroxide* with the fat instead of using sodium hydroxide. The soap is made purer by again dissolving it in water, heating it and once more *salting it out.* This operation is repeated until the soap is made as fine as you want it.

When fine sand or *pumice stone* [1] is added to ordinary soap, *scouring soaps* are formed. Many soaps contain borax and this softens the water without the addition of extra borax or the waste of soap, hence they are good soaps for the laundry.

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[1] *Pumice,* or *pumice stone,* is a porous volcanic lava and is so light it floats on the water. It is used for cleaning and polishing.

Toilet soaps are made in one of two ways, namely, either glycerine or sugar is mixed with the crude soap, when it becomes transparent and beautiful and to make them even more attractive they are colored with dyes and scented with perfumes.

The other way of making toilet soaps is the one in which *ivory soap* is manufactured; that is before the soap solidifies it is beaten by paddles and bubbles of air are thus mixed in it when it is allowed to harden. The air makes it lighter than water and so of course the soap floats.

About Bleaching and Bleaching Compounds.—Bleaching is different from cleaning in that it takes the color out of things whereas the latter only removes the dust and grease. In the chapter on gases you found that chlorine and sulphur dioxide were the great bleaching gases and just what each was used for. You also learned that it was not the chlorine itself but a mixture of chlorine, gas and water which gives *hypochlorous acid* that really did the bleaching and also that sunlight helped the process along.

Now a very common preparation used in the house for bleaching and removing such unsightly things as coffee and tea stains from table linen is known as *Eau de Javelle.*

To make *Eau de Javelle* dissolve about 1 *kilogram* of bleaching powder in about 500 c.c. of water, let it settle and pour off the clear liquor and add to the latter 1 liter of water in which you have dissolved 1 kilogram of washing soda by boiling.

The bleaching agent in this solution is simply hypochlorous acid and for all of its big name it is no more effective than a solution of common chlorine water.

Another makeshift device in use by good housewives for bleaching is the old vinegar salt scheme. The article to be bleached is first dampened and then a paste of vinegar and salt is rubbed into the stain. What really happens here is that the vinegar (which contains acetic acid) acts on the salt and a little chlorine is produced which combines with the moisture in the cloth and forms our old friend hypochlorous acid again and this does the bleaching. As you probably know the sunlight is necessary to make this somewhat ingenious method more effective.

The Action of Bleaching Powder.—While we are still on the subject of bleaching compounds let us examine what is called *bleaching powder*. Bleaching powder is a not very soluble salt whose formula is $CaCl(OCl)$. It is a common staple around the house where it goes under the very familiar name of *chloride of lime*. It is not only a bleaching compound, but a *disinfectant*[1] as well.

The carbon dioxide and moisture of the air have the property of releasing from dry bleaching powder hypochlorous acid; hence the dry powder will disinfect the air and objects in its immediate vacinity.

Further, hypochlorous acid kills bacteria of all kinds and other minute organisms with which sewage is always infected. For these reasons bleaching powder is largely used for all of these purposes.

Lighting by Acetylene Gas.—For use around the house and on small estates where electricity is not to be had and installing such a plant would be too expensive, or require too much care, or both, an *acetylene gas* generator is the next best scheme for lighting. It is shown in Fig. 48.

As a matter of fact, the acetylene gas system of lighting has become quite popular because it is easy to operate, and it makes a very bright, yet soft light.

The gas is generated from *calcium carbide* (CaC_2) by simply adding water to it. The calcium carbide is dis-

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[1] A substance which will kill the germs of contagious diseases such as chlorine, bromine, sulphurous acid, formaldehyde, etc., is called a *disinfectant*.

integrated, that is broken down by the action of the water which drops onto it and acetylene gas is rapidly formed and flows in a steady stream through the pipes. The reaction is as follows:

$$\overset{\text{Calcium carbide}}{CaC_2} + \overset{\text{Water}}{2H_2O} \rightarrow \overset{\text{Calcium hydroxide}}{Ca(OH)_2} + \overset{\text{Acetylene gas}}{C_2H_2}\uparrow$$

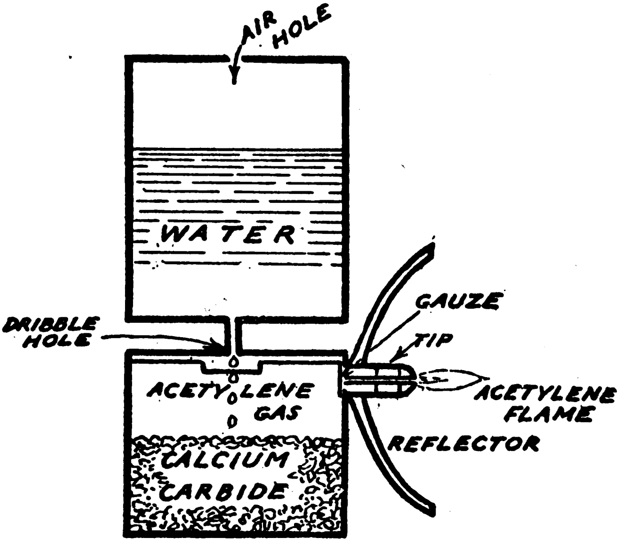

Fig. 48.—How an Acetylene Gaslight Generator is Made

Acetylene gas can be pumped into steel cylinders at ordinary pressures and it can then be safely handled and transported. But when it is under a pressure of over 30 pounds to the square inch the gas explodes at the slightest shock. When pure oxygen is mixed with the acetylene in the right amount, the flame makes an intense heat and this is known as an *oxacetylene flame.* This kind of a flame is widely used for welding metals together and cutting steel and iron girders when buildings are being wrecked.

How to Make Iodine Solution.—A handy preparation to have around the house is a *solution of iodine.* A little of this stuff will cost you real money at a drug store but if you make it yourself it will cost only a trifle.

Get 10 cents worth of *grain alcohol,* take your bottle of *crystals of iodine,* drop them into the alcohol one at a time, and shake until the alcohol takes on a brownish color. You will then have a cure for cuts, sprains, bruises, insect bites and black eyes—should you happen to run into a hitching post.

Of course iodine solution is an external remedy, which means that you must not drink it though its chief ingredient is the same as that which made *Pe-ru-na* famous.

How to Make a Hair Remover.—A superfluous hair remover, or *depilatory paste* as it is called, is usually an expensive article if you buy it at a drug store, and moreover many of them are apt to remove the face and leave the hair. Here is a depilatory paste that is easily and simply made and is not at all harmful. Get a quarter's worth of *calcium sulphide* (CaS) from your druggist and moisten it with water until a rather thick paste results.

Rub this *on* the skin but *not into* it, leave it on over night and it will completely remove objectionable hair. The water acts on the calcium sulphide to form *calcium hydrosulphide* and *calcium hydroxide.* Since alkalies or hydroxides decompose *proteins,*[1] and since hair is really made up of protein this is the substance that does the business.

I might add that calcium sulphide is one of the ingredients of *luminous paint* which has the peculiar property

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[1] The *proteins,* as they are called, are the chief animal and vegetable compounds and they are present in all of the processes of life. They are formed of CH, OH, and S.

of glowing in the dark without giving off heat, but it is not the stuff of which red noses are made.

Lead Compounds Used in Making Paint.—Since lead and its compounds furnish many interesting things that we use in our daily lives, you should know something about them.

When 50 per cent of tin is added to a like amount of pure lead and they are melted together they form a *solder*. Now solder is not only an alloy but it is a *solution* and hence it melts at a much lower temperature than either lead or tin alone.

This makes it possible to apply it to either metal when it is melted without melting them. Lead is used for plumbing because it is easy to bend, can be soldered without trouble and is not attacked by any acid except nitric.

Minum, or *red lead* (Pb_3O_4) is used for making a white paint; by adding colors or dyes to it good oil paints in all colors are made. Since these lead compounds are costly very few ready mixed paints contain them, and therefore when you are painting your house it is the better way to buy the raw materials and either mix them yourself or have a painter actually mix them before your eyes.

Dyes and the Art of Dyeing.—Dyes are substances used for coloring fabrics and it follows that *dyeing* is the process of applying such colors to fabrics.

First get it firmly fixed in your mind that of all fabrics cotton is the hardest to dye fast colors. For this reason stains on wool or silk are often much harder to get out than those on cotton.

Of course the object of dyeing is to color the goods but this must be done in such a way that the dye *is fast*, that is so that it will not fade out by the action of the light or when washed and rubbed with soap.

To get fast colors the kind of dye used and the kind of goods it is used on have much to do with the success of the operation. Cotton and linen are formed of smooth hollow fibers, while wool consists of scaly hollow fibers, and silk is made up of solid filaments as shown at *A*, *B* and *C* in Fig. 49. This is why wool, linen and cotton are so much easier to dye than silk.

There are three kinds of dyes in use and these are (1) *insoluble dyes*, (2) *mordant*, or *adjective dyes*, and (3\ *direct, or substantive dyes*.

FIRST, THE INSOLUBLE DYES.—*Insoluble dyes* are dyes which will not dissolve in water. When this kind of a dye is used the hollow fibers of the materials the goods are made of are filled with a solution of the dye and this precipitates the desired color when brought into contact with some other chemical.

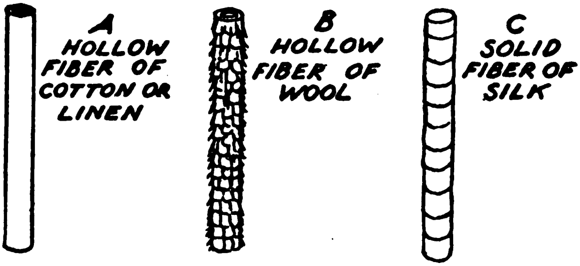

FIG. 49.—DIFFERENT KINDS OF FIBERS

The precipitate is then held in the hollow fibers and since it will not dissolve in water it cannot be washed out. The goods take on the same color as the precipitate in its fibers. While this plan works best on cotton and wool it has been applied to silk, though less successfully.

As an example of this kind of dyeing a piece of goods is first boiled in *lead acetate* $Pb(CO_2CH_3)_2$ which is colorless. It is next soaked in boiling *potassium chromate*

(K_2CrO_2) when it becomes dyed a beautiful permanent yellow. The action is as follows:

$$\overset{\text{Lead acetate}}{Pb(CO_2CH_3)_2} + \overset{\text{Potassium chromate}}{K_2CrO_4} \rightarrow \overset{\text{Potassium acetate}}{2K(CO_2CH_3)} + \overset{\text{Yellow lead chromate}}{PbCrO_4 \downarrow}$$

The lead chromate which is precipitated in the fibers of the goods is yellow and this imparts a brilliant yellow to the piece which is dyed. Since the lead chromate is insoluble, the goods can be washed with soap and water and the color will not run.

Indigo is another insoluble dye and its formula is $C_{16}H_{10}N_2O_2$. It was formerly obtained from Egypt and India but of recent years it has been made *synthetically*[1] in the United States. The goods are first soaked in a solution of *indigo white* ($C_{16}H_{12}N_2O_2$) and the oxygen of the air then oxidizes it, changing the indigo white to indigo blue, whose formula was first written above, thus

$$\overset{\text{Indigo white}}{2C_{16}H_{12}N_2O_2} + \overset{\text{Oxygen and air}}{O_2} \rightarrow \overset{\text{Indigo blue}}{2C_{16}H_{10}N_2O_2 \downarrow} + \overset{\text{Water}}{2H_2O}.$$

NEXT COME THE MORDANT, OR ADJECTIVE DYES.—A *mordant* is a substance for *fixing* a dye. Some goods, such as cotton, are not active chemically and therefore have small tendency to absorb dyes. In order to dye them it is necessary to introduce into the cotton fibers some substance which will absorb the dye first and then carry it into the precipitate and this is what a *mordant* does.

Thus it is with *chromous acetate*, for when a few drops of *sodium hydroxide* are added it causes a precipitate of chromium hydroxide; the chromium hydroxide will then absorb the dye and carry it into the precipitate or goods.

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[1] That is putting different elements together to form a compound exactly like that which nature makes.

A precipitate like this composed of a mordant and absorbed dye is called a *lake*.

Alizarin is a dye which will when added to the above precipitate cause it to turn red, violet, or maroon depending on the amount used. *Cochineal* is another dye used in this way; it is obtained from female insects which live on the *cochineal cactus plant*. Other dyes of a similar kind are *malachite green* and *methylene blue*.

Finally the Direct, or Substantive Dyes.—Dyes made of organic materials are usually fast and direct dyes and need no mordant to make them fast to the material to be dyed. However, such dyes as *Congo red* and *chrysophenin*, after the goods have been soaked in them must be *salted out* by the addition of sodium sulphate in order to coagulate them, just as sodium chloride is added to a soap solution to coagulate the soap.

How to Waterproof Goods.—It is sometimes handy to know how to *waterproof* goods. This is done by first soaking them in *aluminum acetate* and then *steaming them*. The steam *hydrolizes* the aluminum acetate and *aluminum hydroxide* is produced and precipitated.

It fills the hollow fibers of cotton, linen or wool and makes them non-absorbent and impervious to moisture. The action is as follows:

$$\overset{\substack{\text{Aluminium}\\\text{acetate}}}{Al(CO_2CH_3)} + \overset{\text{Steam}}{3H_2} \rightarrow \overset{\substack{\text{Aliminium}\\\text{hydroxide}}}{Al(OH)_3\downarrow} + \overset{\text{Acetic acid}}{3HCO_2CH_3}.$$

How to Fireproof Goods.—Here is a way to fireproof goods that makes them about as proof against fire as a fireproof hotel and you know how much that is. This fireproofing compound is called *water glass* ($Na_2Si_2O_5$) and when a fabric is soaked in it or it is brushed over wood it fills the pores of the material with a hard, glassy layer which is quite invisible. It is made by dissolving *silica*, that is *silicon dioxide* (SiO_2) with *caustic soda*

(NaOH) in boiling water when *sodium silicate,* that is *water glass* is formed thus:

$$\overset{\substack{\text{Silicon}\\\text{dioxide}}}{SiO_2} + \overset{\substack{\text{Caustic}\\\text{soda}}}{2NaOH} \rightarrow \overset{\substack{\text{Water glass}\\\text{Sodium silicate}}}{Na2SiO_3} + \overset{\text{Water}}{H_2O}$$

You know that sand will not burn for the reason that it is silicon dioxide (SiO_2) and contains as much oxygen as it is possible for the silicon to combine with and as it can't combine with more oxygen it cannot burn for burning consists of something combining rapidly with oxygen.

CHAPTER XI

THE ELECTRO-CHEMICAL ARTS

How to Generate an Electric Current
Kinds of Battery Cells
The Parts of a Primary Cell
How a Dry Cell is Made
    What Manganese Dioxide Does
The Gravity, or Crowfoot Cell
The Bunsen Cell
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The Electrolysis of Liquids
How to Decompose Water by Electricity
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CHAPTER XI

THE ELECTRO-CHEMICAL ARTS

As you have already guessed from the heading of this chapter *electro-chemistry* has to do with *electricity* acting in combination with *chemical* elements and compounds.

Stated a little more in detail electro-chemistry is based on either (1) the production of an electric current by the chemical action of certain elements such as the metals, upon compounds, such as acids and the salts of metals, or (2) the action of an electric current on elements such as metals and certain compounds when these substances are disintegrated and decomposed.

How to Generate an Electric Current.—The first thing you need to perform any kind of an electro-chemical experiment is an *electric cell*, or as it is more often, though wrongly called a *battery* for it is only when two or more cells are coupled together that a battery is formed.

Now the purpose of a cell is to generate a current of electricity and a cell of whatever kind gets its power to generate a current from the electro-chemical action of the chemicals of which it is made.

Kinds of Battery Cells.—There are two distinct kinds of *electric* cells and these are (1) the *primary cell*, and (2) the *secondary cell*, or *storage battery* as it is commonly called.

There are several kinds of primary cells but as all of them are made and work on the same principle I shall tell you about just three of them and these are (1) the *dry cell*, (2) the *gravity*, or *crowfoot cell* and (3) the *Bunsen cell.*

The Parts of a Primary Cell.—A primary cell of any kind is made up of (1) the *jar* that holds the solution and the elements, (2) the *solution*, or *electrolyte* as it is called, (3) the *cathode*, or *negative element* which is best made of zinc, and (4) the *anode*, or *positive element* which is usually of *carbon*, as shown in Fig. 50.

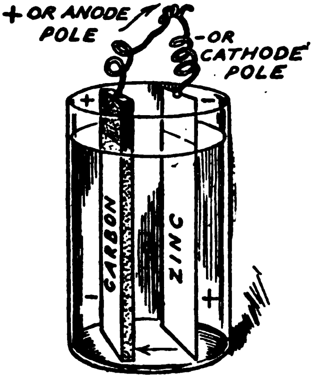

FIG. 50.—AN ELECTRIC CELL

How a Dry Cell is Made.—The *dry cell* is the most common kind and millions of them are used for bell, telephone and ignition work. A dry cell is not, strictly speaking, *dry* for the chemicals it contains are wet up with water.

The cell is formed of a cup of sheet zinc which serves as the jar and for the negative element at the same time; to the top of this cup a binding post is fixed as shown in Fig. 51. The *active* paste, or *electrolyte* as it is called, is made of 1 part of *zinc oxide*, 1 part of *zinc chloride*, 1 part of *ammonium chloride*, whose common name is *sal ammoniac*, 2 parts of *manganese dioxide*, 3 parts of *plaster of Paris*, mixed well with 1 part of water, all by weight.

The zinc cup is filled to within ½ an inch of the top and a carbon rod, for the positive element, with a binding post fixed on one end of it is pushed down into the paste to within an inch of the bottom. Melted pitch is then poured over the paste flush with the top of the cup when the cell is ready to use. The chemicals in this cell do not act on the zinc except when the circuit is *closed*, that is when the zinc and carbon are connected together.

What the Manganese Dioxide Does.—A little further on I will have a few words to say about how a cell generates a current but right here I will tell you why the *manganese dioxide* is used.

When a cell is delivering current bubbles of hydrogen are set free from the zinc and if allowed to they are carried to the carbon and soon cover it. This action is called *polarization* and it tends to cut down the amount of current the cell is giving. To prevent this untoward action manganese dioxide is mixed with the paste and this slowly gives off oxygen which combines with the hydrogen before it is carried to the carbon and in this way all of the current can pass.

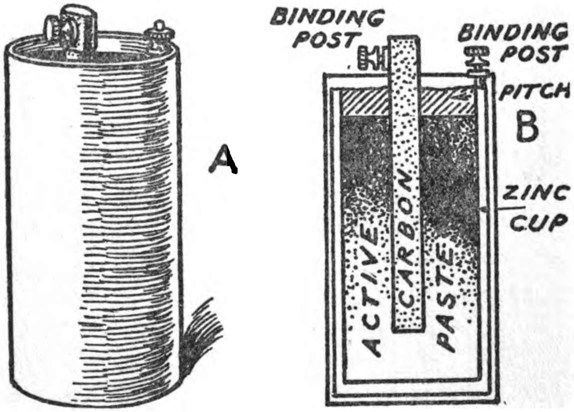

Fig. 51.—How a Dry Cell is Made

The Gravity, or Crowfoot Cell.—This kind of a cell, which is shown in Fig. 52, is largely used for working telegraph systems and is what is known as a *closed circuit* cell, that is the elements of the cell must be kept connected together or the solution will spoil.

The positive element is formed of a thin strip of copper bent in the shape of a star and this sets on the bottom of the jar; an insulated wire is soldered to it and is lead to a binding post on the top of the jar.

About half a kilogram of *copper sulphate,* or *blue vitriol* as it is called, is placed around the copper star and 250 c.c. of water is poured in in which a few drops of sulphuric acid have been stirred. The negative element is made of zinc cast in the shape of a crow's foot, and it is from this peculiarity that the cell gets its name; the zinc is hung on the edge of the jar into the solution when the battery is ready for work.

FIG. 52.—A GRAVITY, OR CROWFOOT CELL

The cell is called a *gravity* cell because the sulphuric acid acting on the zinc forms *zinc sulphate* and as this is lighter than the *copper sulphate* the zinc sulphate solution floats on the copper sulphate solution.

To keep this cell in good condition both elements must always be kept connected otherwise the copper sulphate will build up a crust on the zinc crowfoot and prevent any further action from taking place.

The Bunsen Cell.—While the dry cell and gravity cell can be used for electro-plating on a small scale, a *Bunsen cell* is the right kind to use for this purpose. The

Bunsen cell, as shown in Fig. 53, is different from those described in that two different liquids are used for the electrolytes and these are separated by a *porous* cup.

The jar is nearly filled with dilute sulphuric acid and in this is set a cylinder of zinc; the porous cup is set inside the zinc cylinder, this is nearly filled with concentrated nitric acid and into this a carbon rod is put.

The hydrogen set free by the action of the sulphuric acid on the zinc passes through the porous cup and acting on the nitric acid changes it to *nitrogen peroxide* (NO_2) which escapes as brownish-red fumes. To make the cell work its very best it should be cleaned and the electrolyte renewed every day.

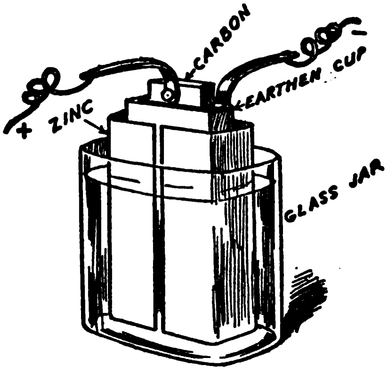

Fig. 53.—A Bunsen Cell

How a Cell Gives an Electric Current.—You have already seen that when a chemical change takes place some form of energy is set up and this is usually heat. But if the right elements and chemical compounds are used as in a battery cell not only will heat be given off but a current of electricity will be developed as well.

It is not my idea to tell you just how a cell generates a current for the chemical actions set up are very complex and hence difficult to grasp but it will be enough to say that the atoms of all the elements are made up of little charges of electricity called *electrons* and when a metal like zinc is acted on by an acid, or an alkaline solution, these minute electric charges are set free and flow in a steady stream, or *current*, from the negative element to the positive element through the electrolyte when they are conducted through the outside circuit and thence back from the positive to the negative element.

The Secondary, or Storage Cell.—*How it is Made.*—Storage batteries, which are made up of a number of *secondary cells*, are largely used for running electric automobiles, in electric power plants, for telegraph and telephone work, etc.

A storage cell does not generate a current like a primary cell by direct chemical action but it must be *charged*, as it is called, by a current of electricity first and when this is done it will deliver a current until the cell is discharged.

A storage cell (see Fig. 54) consists of a jar containing a dilute solution of sulphuric acid and in this is immersed a number of negative and positive elements. Different from those of a primary cell, these elements are made of lead plates, or *grids*, as they are called, and these have holes or grooves in them which are filled with an active material.

The holes, or grooves in the negative plate are filled with finely divided lead (Pb) which makes it active and the positive plate is filled with *lead dioxide* (PbO_2) which is the opposite active material. All of the negative plates are connected together and all of the positive plates are connected together, while all of them are kept apart by thin boards, or *separators* as they are called.

HOW A STORAGE CELL WORKS.—Let us see, now, what action goes on when a storage cell is discharging, that is delivering a current. First the *negative all lead plate* and the *positive lead dioxide plate* are acted on by the sulphuric acid and a coating of *lead sulphate* is formed on both plates, and at the same time water is also formed.

When the cell is being charged by an electric current the water acts on the lead sulphate and this removes the films of this substance from the plates thus leaving them as they were at first, when sulphuric acid is formed. The formula which follows shows how these actions take place:

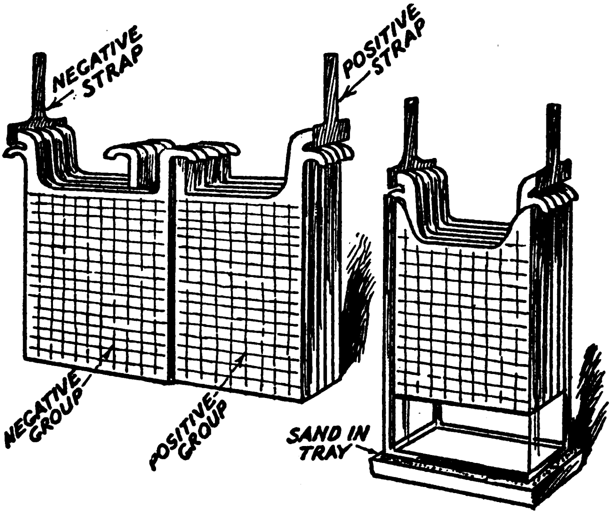

FIG. 54.—HOW A STORAGE BATTERY IS MADE

← CHARGE

Lead	Sulphuric acid	Lead dioxide	Lead sulphate	Water

$$Pb + 2H_2SO_4 + PbO_2 \leftrightarrows 2PbSO_4 + 2H_2O$$

DISCHARGE →

The Electrolysis of Liquids. By the *electrolysis of liquids* is meant the *decomposing,* that is the breaking up of liquid compounds by passing a current of electricity through them.

Now if you will think back to the second chapter you will recall that water is made up of 2 parts of hydrogen and 1 part of oxygen and that I showed you how to make water by burning hydrogen in air, the air furnishing the necessary oxygen, but this experiment did not show you in just what proportions they combine.

It is my purpose now to tell you how you can learn the proportions after breaking down the water by means of passing an electric current through it, this action causing it to separate into its two original elements, namely hydrogen and oxygen. From the amount of each of these two gases formed in this way you will be able to see exactly in what proportions they combine.

How to Decompose Water by Electricity.—Get a shallow glass dish and pour into it a dilute solution of sulphuric acid made as follows: into 100 c.c. of water *stir slowly* 10 c.c. of concentrated sulphuric acid.

Never stir water into sulphuric acid since much heat is developed and may turn the water into steam, in which case it will throw the acid out and cause it to splash on you. Besides the acid is heavier than water and so it will mix easier when poured into the latter instead of doing it the other way.

Fill two test tubes, with the rest of this solution and turn them upside down in the dish taking care not to let any air get into them. Next get a couple of dry cells and connect the carbon of one to the zinc binding post of the other with a short piece of copper wire, and attach a long piece of copper wire to each of the other binding posts of the battery. Twist the free ends of each of these wires to a bit of carbon 2 inches long and put it on tight, but first scrape the wire clean with your knife.

Now slip one of the carbons into each test tube taking care not to let any air enter and the apparatus is now ready for the electrolysis of water as shown in Fig. 55.

You will note that bubbles of gas immediately begin to form at the carbons and rise to the tops of the test tubes driving the water out of them. Should these bubbles form very slowly, or not at all, it may be necessary to connect a couple of more dry cells in the circuit, or to increase the amount of sulphuric acid in the solution which fills the dish and the test tubes.

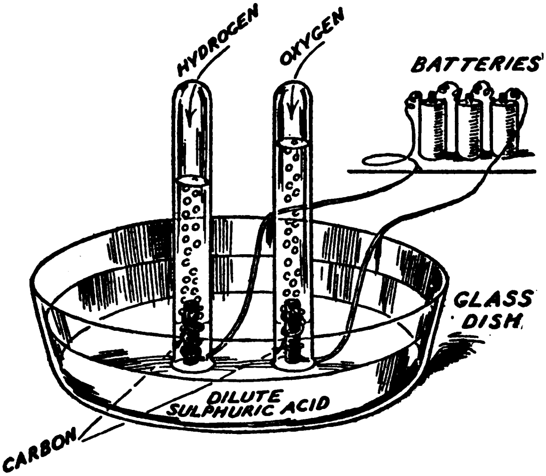

FIG. 55.—THE DECOMPOSITION OR BREAKING DOWN OF WATER BY ELECTRICITY

You will notice that in one of these tubes the water is driven out *twice as fast* as from the other tube. This is the tube in which hydrogen gas is being filled, and since the water which is being broken down contains twice as much hydrogen by volume as it does oxygen, it is quite natural for the water to be driven out of the tube twice as fast as is the water in the oxygen tube.

To prove that this tube contains hydrogen, wait until all of the water has been driven out and then hold a lighted match, or other flame close to the mouth of the tube. It burns quickly and makes a sharp report and this proves that it is hydrogen. When the water has been displaced by the gas in the other tube, put a glowing match into it, when it will burst into a flame and this proves that the tube contains oxygen.

This means that the current of electricity has decomposed or broken down the water into its two original gases. Now water in itself will not *conduct*, that is carry, a current of electricity. There are, however, certain chemical compounds which when added to water are dissolved by the water and makes it a fairly good conductor of clectricity. The chief compounds and substances which have this property are acids and salts. This is the reason you added the sulphuric acid to the water instead of trying to decompose the water without it. This method of decomposing water by passing an electric current through it is known as the *electrolysis of water*.

How Acids Conduct Electricity.—Before I go on and explain further how this action works I must tell you a little more about acids and salts. I have already told you that acids always contain hydrogen, which is known as the *acid radical*, in combination with some other element.

I have also told you that when acids that are free from water do not have any of the same chemical properties that they do when in solution, that is mixed with water, and this is likewise true of their conducting electricity. However, as soon as they are dissolved in water, both the acids themselves and the solution becomes a good conductor of electricity.

The reason is this: When an acid is dissolved in water the acid radical, or hydrogen, takes on a positive charge of electricity which is written +, and the other element, or radical, which it is combined with, takes on a negative charge of electricity, and is written −.

Thus when hydrochloric acid is added to water the hydrogen takes on a positive charge of electricity and the chlorine takes on a negative charge, thus:

$$\overset{\text{Hydrochloric acid}}{HCl} + \overset{\text{Water}}{HO_2} \rightarrow \overset{\text{Hydrogen positively charged}}{H^+} + \overset{\text{Chlorine negatively charged}}{Cl^-}$$

Moreover the acid no longer acts simply as a compound but the electrically charged hydrogen and chlorine act as separate and distinct substances.

Now, when the two terminals or ends of the wires of a battery are placed in this solution the electric charges flow through it from one terminal to the other, hence one of these terminals is known as the positive or + terminal and the other as the negative, or − terminal.

It is a well-known fact that positive and negative electricity always attract each other, consequently when the terminals of a battery are dipped into hydrochloric acid, the hydrogen which is *positively charged* goes to the *negative pole* of the battery and escapes there as a gas. The chlorine which is *negatively charged* is attracted to the *positive pole* of the battery and also escapes as a gas. This is the way hydrochloric acid is broken down and the process is known as the *electrolysis of hydrochloric acid.*

As soon as the hydrogen reaches the negative terminal of the battery it loses its positive charge and likewise the chlorine on reaching the positive terminal loses its negative charge. Hence these gases are able to escape as gases, while if they still held their charges they could not do so.

If the products of the decomposition of a substance by electrolysis are very soluble in water they will dissolve and will not escape as a gas.

How Water is Decomposed by Electricity.—In the electrolysis of water the current of electricity acts on both the water and on the sulphuric acid which you added to it in order to make the water a conductor. It breaks the sulphuric acid down and it breaks the water down into hydrogen and oxygen. The decomposed sulphuric acid is made up of products which are very soluble in water and consequently they dissolve in the water to form more sulphuric acid. Hence this part of the reaction is disregarded since the only noticeable products given off are the hydrogen and oxygen of the water. The action is as follows:

$$H_2O + \text{electricity} \rightarrow \overset{\substack{\text{Cathode or}\\ \text{negative pole}}}{2H^+\uparrow} + \overset{\substack{\text{Anode or}\\ \text{positive pole}}}{O^-\uparrow}$$

About Electroplating.—The electrolysis of water, however, is the least important use to which this property of electricity to break down acid and salt solutions can be put. Since the products of electrolysis are deposited at the pole of a battery the art of what is known as the *electro-decomposition of metals*, or *electroplating*, as it is commonly called, is made possible.

This branch of electrochemistry is extremely valuable for metals which rust can be plated with non-rusting metals and the inferior metals can be coated with precious metals, etc.

Thus, if you have a knowledge of electroplating you can do many interesting and useful things such as plating, that is, coating, old spoons and other silverware with silver and making them look like new, plating brass and iron house-fixtures such as towel racks, etc., with nickel, and plating jewelry made of cheap alloys or common metals with silver and gold.

What You Need for Electroplating.—I shall not try to tell you the exact chemistry of all the operations used in electroplating because they are often very complex and to grasp the meaning of them you would need a more complete knowledge of chemistry then I have given in this book.

However, I shall tell you all you need to know about electroplating so that you can do it and then something about the reactions which takes place during the process.

To do electroplating of any kind you need (1) a couple of dry, or gravity cells or better, a single Bunsen cell, (2) the plating solution, or *bath* as it is called, (3) a piece of metal of the kind you are going to plate with and which is called the *anode*, and (4) the article which you want to plate. And now to do a little electroplating.[1]

How to Clean Objects to be Plated.—To plate one metal with another metal successfully you must have the object to be plated perfectly clean and the way to do this is to put it in a boiling-hot solution of *potash*, that is *potassium carbonate* and water until the grease is cut off.

If the object is tarnished it must be cleaned with a *brass scratch brush* first and then dipped into the potash solution. This done rinse the object in clean water to remove the potash and then put it in the plating bath. Do not touch the article after it is clean with your fingers.

How to Plate with Copper.—One of the easiest things in the world to do is to electroplate with copper, for all you need is to connect a piece of copper for the anode to the *carbon*, or + pole of your battery with a wire, and a silver coin with the zinc, or – pole and suspend both the copper and the coin in a strong solution of *copper sulphate*, that is blue vitriol, as shown in Fig. 56.

The copper sulphate is decomposed by the current,

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[1] All kinds of electroplating apparatus and supplies are sold by Hanson & Van Winkle, Newark, N. J.

the copper from the salt will be deposited on the coin, or + electrode, and the sulphuric acid will go to the plate of copper or + electrode, which is called the *anode*. Since the cheaper metal is deposited on the more valuable one plating a coin with copper is of no value except as an experiment. But electroplating with copper is largely used in a process for making a metal copy of a cut, of type or the like, for printing; this is called *electrotyping* and I shall describe this process presently

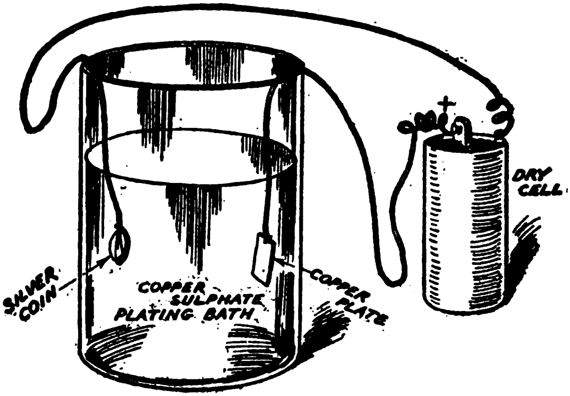

Fig. 56.—Plating a Coin with Copper

How to Plate Iron and Steel with Copper.—You may some time want to plate iron or steel articles with copper and you can do so if you will carefully follow out these instructions.

Take the iron or steel article and scrape it until it has a smooth, bright surface. Then dip it into *benzoline*, *naphtha*, or *gasoline*, to cut off whatever grease there might be on it. Now make a solution of 90 c.c. of water and 10 c.c. of sulphuric acid, and dip the article into it. Rinse it off with clear, cold water and it is then ready for the plating bath.

Fasten the article to the zinc of the battery and a plate of copper to the carbon of the battery, then hang them in a bath which you have made by dissolving 40 grams of *double cyanide of copper and potassium,* 10 grams of *concentrated ammonia,* and 10 grams of *potassium cyanide.*[1] When the article has been well plated remove it from the bath, wash it off in clean, cold water and dry it.

How to Plate with Nickel.—Many articles of iron around the house which are near water tend to rust and this can be prevented by nickel plating them. Besides nickel will take on a high polish and iron and brass articles always look much better when nickel plated.

To plate with nickel you must first clean the objects as just described under copper plating. For the bath make a solution as follows: In 1000 c.c. of water dissolve 50 grams of *double sulphate of nickel and ammonium* and 25 c.c. of *ammonium sulphate.*

Use a nickel piece for the anode and the article to be plated for the cathode.

The nickel anode must be as well cleaned as the object to be plated, and when once clean see to it that your fingers do not touch it, for wherever they do a film of grease will be left and the plating will not stick to the surface.

Lower the nickel anode and the article to be plated into the bath and let the plating process go on until the nickel is as thick as you want it on the article.

If you want a high polish on the nickel coat it with a mixture of *rottenstone*[2] and machine oil. Use a flannel rag and plenty of elbow grease and rub the article until you can see yourself in it. A better way, because it is easier, is to buff it on a fast rotating felt wheel using rottenstone as before and then polishing it with a rag wheel.

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[1] These are very poisonous chemicals.

[2] A soft kind of limestone.

How to Plate with Silver.—To be able to plate with silver and gold will prove a good deal of service to you if you care for the money end of it, for nearly everybody has some trinket or other which they want plated or some plated table wear which they want replated. Moreover, silver and gold plating is much easier to do than nickel plating.

To plate with silver scrape the article or object bright, and clean it as I explained under the heading of *How to Clean Articles to be Plated.* Make a plating solution by dissolving 10 grams of *silver cyanide* (SCN) and 15 grams of *potassium cyanide* (KCN) in 1 liter of water and hang the anode and the article to be plated in it as shown in Fig. 57 or, better, make a regular suspension as shown in Fig. 58.

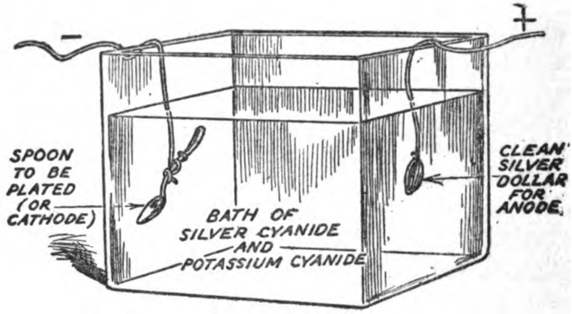

Fig. 57.—How Silver Plating is Done

Now dip the article which has been cleaned into a bath made by dissolving 10 grams of *mercuric nitrate* ($Hg(NO_3)_2$) in 1000 c.c. of water, leave it there for a second and then place it in the plating bath. The mercuric nitrate will give it a thin coating of mercury and the silver will deposit on this very easily.

Burnishing Silver-plated Ware.—To give a high finish to silver- and gold-plated articles they are not buffed but *burnished.* A *burnisher* is a tool made in various shapes of polished steel or agate; the burnisher is dipped in soapy water and rubbed to and fro over the part to be burnished with considerable pressure. The result is a bright, smooth surface produced such as you see on silver and silver-plated ware.

How to Plate with Gold.—To plate jewelry and other small articles with gold, clean them thoroughly. A little piece of gold about the size of a gold dollar must be used for the anode and make the plating solution as follows:

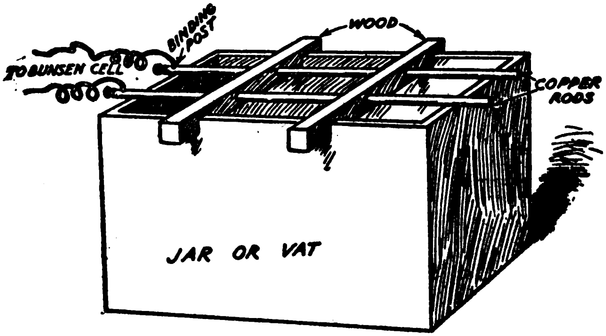

FIG. 58.—HOW A REGULAR PLATING VAT IS SET UP

Dissolve 10 grams of *gold chloride* and 100 grams of *potassium cyanide* in 1000 c.c. of water. The article to be plated should then be *quicked,* that is dipped in the *mercuric nitrate* as described for silver plating.

Connect the article to be plated to the zinc pole of the battery and suspend it in the gold bath until enough gold has been deposited on it to give it a 14-carat fine appearance. Then take it out, wash it, dry it in boxwood sawdust and burnish it.

How to Make Electrotypes.—One of the most important processes in which electroplating with copper is used is in making *electrotypes*. This process consists of making a mold or *matrix* as it is called, of plaster of Paris, or other materials, of the type as shown at *A* in Fig. 59 and plating the mold with copper to form a shell as shown at *B*.

The plated shell of copper is then backed up with type-metal to a thickness of about $\frac{3}{16}$ of an inch; it is then mounted on a block of wood which makes the whole electrotype just the height of ordinary type, when it is ready to be printed from.

The way in which an electrotype is made is like this: After the plaster of Paris mold, or matrix, of the type has been made it is dipped into melted paraffin so that it will not go to pieces when it is put into the plating bath. Since both plaster of Paris and paraffin are non-conductors of electricity, the mold, or matrix, must be well brushed with powdered *graphite*,[1] or *blacklead* as it is sometimes called, and this makes it a very good conductor.

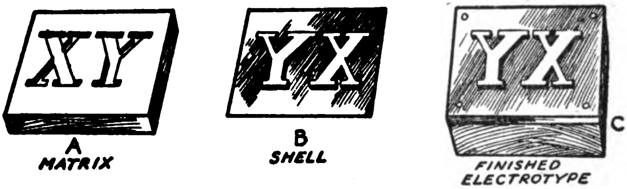

FIG. 59.—MAKING AN ELECTROTYPE

The mold, or matrix, is then attached to the zinc of the battery by means of a wire twisted around it, the copper anode is connected to the carbon of the battery and both the matrix and anode are then hung in the solution, or bath.

This is made by dissolving 1½ kilograms of *crystallized copper sulphate* in 1 liter of water and adding to it 1 kilo-

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[1] A very soft flaky kind of carbon.

gram of sulphuric acid The mold is kept in the plating bath until the film of copper formed on it is thick enough and then the type metal is poured on top of and into it; when cold the electrotype can be taken from the mold and mounted on a block of wood as shown at *C*.

Electroplating Simply Explained.—As you have had occasion to observe the plating solution, or bath, consists of a salt dissolved in water. Further, as you know, a salt is made up of a metal and a non-metal, as for instance *sodium chloride* which is formed of the metal sodium and the non-metal chlorine.

When a current of electricity is passed through this solution the latter is decomposed. The metal of the salt which has a positive charge goes to the *cathode*, that is the negative, or zinc pole of the battery, and the non-metal which is negatively charged, goes to the *anode*, or positive or carbon pole of the battery.

The cathode pole, however, is attached to the article you want to plate with a wire and this soon becomes plated with the *metal of the salt*, while the anode, or metallic plate becomes covered with the non-metal of the salt. However, as the metal of the salt is used up in plating the article it is gradually replaced by the metal from the plate which goes into the solution and finally the plate of metal will entirely disappear.

To understand this fully look at the following equation:

$$\text{Salt}+\text{electricity}\rightarrow\underset{\substack{\downarrow\\ \text{goes to object}\\ \text{to be plated (cathode}-)}}{\text{metal (of salt)}^{+}}+\underset{\substack{\downarrow\\ \text{goes to the}\\ \text{metallic plate (anode}+)}}{\text{non-metal (of salt)}^{-}}$$

From this you can get a pretty clear idea of how electroplating is carried on. By using different kinds of salts and anodes of different metals it is possible to plate with almost any metal, though some are harder to deposit than others.

CHAPTER XII

THE CHEMISTRY OF PHOTOGRAPHY

WHAT LIGHT DOES

WHAT LIGHT IS

THE ACTION OF LIGHT ON SILVER CHLORIDE
- How to Make Silver Chloride

PHOTOGRAPHY SIMPLY EXPLAINED

HOW DRY PLATES ARE MADE

HOW A DRY PLATE TAKES A PICTURE
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- Why a Fixing Bath is Needed

ABOUT PRINTING FROM NEGATIVES

WHY FILMS ARE USED

KINDS OF PRINTING PAPERS
- Velox Paper
- Silver and Solio Papers

WHY SILVER PRINTS ARE TONED

HOW TO MAKE BLUE PRINTS
- How to Write on Blue Prints

CHAPTER XII

THE CHEMISTRY OF PHOTOGRAPHY

The next time you go to the *movies* just stop looking at the picture long enough to think that if it was not for chemistry you'd be at home studying your arithmetic like the fellows of twenty years ago. So be glad you are living *now* and not *then*. And when I tell you that a cat was responsible for the discovery of modern photography, I want you to be as surprised as was his master. So read on.

What Light Does.—It may seem strange on first thought that light will act on certain chemical compounds and break them down very much as heat does in *effect*, though the *action* of light is, of course, very different from that of heat.

It is due to this remarkable property of light that the *science of photography* has been given us, from the simple *blue print* to the complex *motion picture film* and other wonderful effects that were undreamed of by those who lived a short century ago.

What Light Is.—If you drop a stone into a pond little *waves* of water will form in circles around it and spread out to the shore.

When you strike a bell the rim of the bell moves rapidly forth and back, as shown by the broken lines in Fig. 60, or *vibrates*, as it is called; these movements disturb the air and send out *waves* in it, which we call *sound waves*, and when they reach your ear you *hear* the bell ring.

In exactly the same way when anything burns and makes a light, the atoms of burning gas are set into very rapid motion, or *vibration* as it is called and these send out *waves* in the *ether*.[1] When these *waves*, which we call *light waves*, impinge on your eyes the latter are affected and you *see* the thing that is burning. (See Fig. 61.)

Now the *note* the bell rings out depends on the number of times it vibrates in a second, that is when it vibrates rapidly it sends out *short waves* through the air and the note you hear is high. The slower the bell vibrates the longer will be the waves in the air and hence the deeper the note.

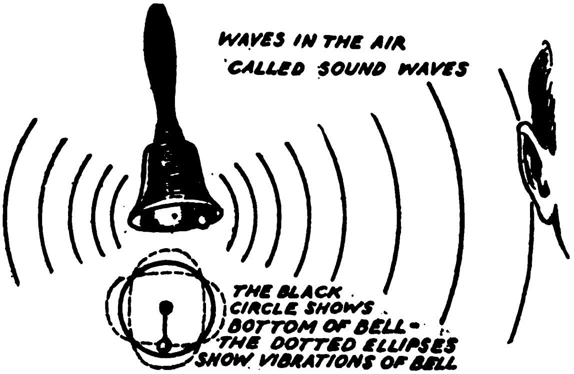

FIG. 60.—HOW SOUND WAVES ARE MADE

Just so it is with a light, that is, the more rapid the vibration of the atoms of gas producing it the *shorter* will be the light waves sent out by them through the *ether*. When these waves are very short they make us see the *color* we call violet, but as the vibrations of the light get slower, the wave lengths in the *ether* get longer and when these are extra long we see the color *red*. Other wave lengths in between the very short and extra long give the other colors of the rainbow.

Now the sun, the electric light, burning magnesium,

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[1] A highly elastic and transparent kind of matter that fills all space and penetrates the densest metals. It is the medium through which light, magnetism and electricity travel.

etc., all send out waves of every length but it is only the shorter wave lengths which make violet light, and it is these short wave lengths, or *violet light* as we say, that act strongly on chemical compounds, and the longer they are the less effect they have, until the extra long wave lengths which make red light have little or no effect at all on chemical compounds.

This is the reason, then, why a red light can be safely used for developing a photographic *dry plate*, or *film*, or a *velox print*. This is the *physics* of light and now let us look into the *chemistry* of light.

The Action of Light on Silver Chloride.—A very common compound which light acts on and which is largely used in photography is *silver chloride* (AgCl). You can easily make some and see for yourself just what happens when it is exposed to the light.

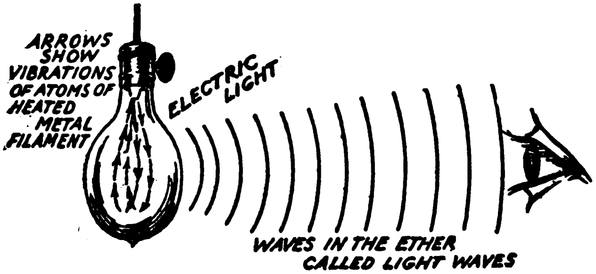

FIG. 61.—HOW LIGHT WAVES ARE MADE

HOW TO MAKE SILVER CHLORIDE.—Dissolve 5 grams of *sodium chloride*, that is, common table salt, in about 15 c.c. of water and 5 grams of *silver nitrate*, which is a nitrogen salt of silver, in about 5 c.c. of water.

When both of the salts have been dissolved pour the sodium chloride solution into the silver nitrate solution and you will see that a milky white precipitate is formed. This is *silver chloride* and the action is shown in Fig. 62.

Now fold a piece of filter paper and fit it into your glass funnel, pour the contents of the test tube containing the precipitate and the solution through it. The precipitate will remain on the filter paper and the liquid will go through it. The liquid is *sodium nitrate* ($NaNO_3$), and the action is as follows:

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

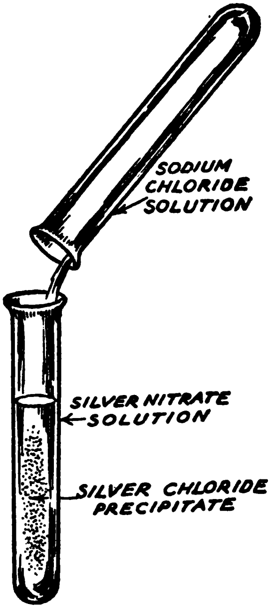

FIG. 62.—HOW SILVER CHLORIDE IS MADE

This is clearly a double decomposition; since the AgCl will not dissolve, it remains in the solution as a precipitate while the $NaNO_3$ being soluble dissolves. Take the filter paper with the AgCl on it and stand it in the bright sunlight.

You will observe that it soon gets dark, turning a purple color first and then brown. What has really happened is that the light has broken down the silver chloride into its original elements, that is silver and chlorine thus:

$$\overset{\substack{\text{Silver}\\\text{chloride}}}{AgCl} + \text{light} \rightarrow \overset{\text{Silver}}{Ag} + \overset{\text{Chlorine}}{Cl}\uparrow$$

The chlorine goes off as a gas and the silver powder that remains is composed of very fine particles of *pure silver*, which have a brownish look.

Now this action, and many others like it, are the greatest factors in photography for by means of them we are able to get the picture of an object on a plate of glass, or a celluloid film or a piece of paper just as it looks in real life, minus the color.

If you will take a careful look at a photograph you will notice that it consists of a lot of light spots and dark patches the whole scheme making up the picture. While a white and black picture is enough for us to "see ourselves as others see us," it is after all only a record of light and shadow.

One's eyes are different from a photographic plate in that the former have the peculiar property of *color sensation,* that is different wave lengths make us see the different colors, while in an ordinary photograph the colors are not present and the picture is simply one in white and black.

Photography Simply Explained.—Since you know how light acts on *silver chloride* it must be clear that if you could coat a sheet of glass with this substance in a dark room and then let an image be projected upon it, that is reflected from some person or object, the light where it fell on the plate would act on the silver chloride as described above.

Also, when it did this it would set the chlorine of the gas free and leave the brown powder of *silver behind on the plate.* Further, since some colors reflect more light then others there will be more of the brown silver chloride formed in some places on the plate than on others.

Thus if you exposed the plate to a block of wood half of which was painted white and the other half painted black, you would find that the part of your plate which the white had acted on would have a large amount of brown silver on it, in fact, so much so that it would look black, while the part that the black acted on would not have changed at all, that is to say, it would still be white. A picture of a negative and a print made from it is shown in Fig. 63.

All of this would take place because white reflects all of the light that strikes it and so changes all of the silver chloride while black absorbs all of the light and so does not reflect any of it when the silver chloride remains unchanged.

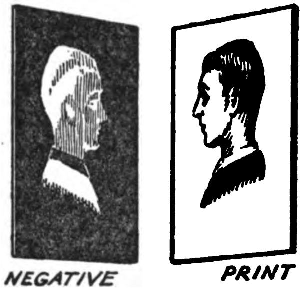

FIG. 63.—A PICTURE OF A NEGATIVE AND PRINT

And so on with the other colors such as blue, green, yellow and red; all of these colors give off more or less light as the case may be and when they act on the plate they will cause either a large or small amount of pure silver to be formed thus making the light and dark patches that appear on the plate.

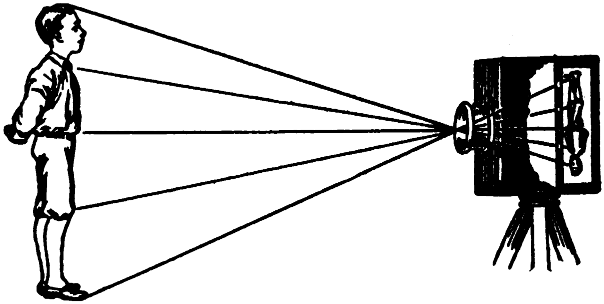

FIG. 64.—HOW A CAMERA ACTS

The practical part of a photographic outfit consists of a *camera* and a *dark room*. The camera (see Fig. 64) is formed of (1) a *lens* which projects the image of the person or object on the plate, (2) the *plate holder* which holds the sensitized plate and (3) the *camera box* which holds the lens and the plate holder in position and keeps all of the light away from the plate except that produced by the image. How *dry plates*, as the photographic plates are called, are made, developed and fixed will be told as we go along.

How Dry Plates are Made.—The silver coating on dry plates and films for cameras must be as sensitive as possible so that the instant the light strikes the sensitized surface the silver compound will be changed and, which means in this case, that the picture is made. I believe I told you before that there are other substances besides silver chloride which light acts on and the one which is used chiefly in photography is *silver bromide*. This is used because light acts on it much quicker and better than it does on silver chloride, or to put it another way, silver bromide is more *sensitive* to light than silver chloride is.

Now silver chloride and silver bromide are compounds which are as like as two peas, and so what I have told you of the one will be about enough for the other. The glass plates for the dry plates are cut from sheet glass to the required size.

What is known as an *emulsion* to sensitize the plates with is then prepared as follows: *silver nitrate*, *ammonium bromide* and *gelatine* are heated slowly together until the reaction has taken place between the first two named substances and *silver bromide* is left thus:

$$\overset{\substack{\text{Silver}\\\text{nitrate}}}{AgNO_3} + \overset{\substack{\text{Ammonium}\\\text{bromide}}}{NH_4Br} \rightarrow \overset{\substack{\text{Ammonium}\\\text{nitrate}}}{NH_4NO_3} + \overset{\substack{\text{Silver}\\\text{bromide}}}{AgBr}$$

As you can plainly see this reaction is a double decomposition. The silver bromide collects in little drops, or globules, in the gelatine and this process is known as *ripening*. Then the emulsion, as the mixture is now called, is allowed to cool and when it gets about as thick as jelly it is cut into little square pieces.

Next it is soaked for a long time in water, to dissolve the ammonium nitrate out of the gelatine but the silver bromide is left in it. The gelatine containing the silver bromide alone is then melted again and one side of the glass plate is dipped into it.

The greatest care must be taken to get the film of gelatine on the glass of the same thickness all over[1] and it is then allowed to harden when the dry plate is made and ready to use.

How a Dry Plate Takes a Picture.—When the plate is put in the camera and the *shutter* of the lens is opened the light from the object to be photographed passes through the lens and strikes the silver bromide which it decomposes just the same as it does silver chloride thus:

$$\overset{\substack{\text{Silver}\\\text{bromide}}}{AgBr} + \text{light} \rightarrow \overset{\substack{\text{Pure}\\\text{silver}}}{Ag} + \overset{\text{Bromide}}{Br}\uparrow$$

leaving metallic silver which is in the form of a brown powder.

Now since the dry plate is only *exposed* to the light for a very small fraction of a second this change in the silver bromide cannot be seen and the picture will not appear until another operation called *developing* has been gone through with, for the change of silver bromide to metallic silver has not been completed.

How a Dry Plate is Developed.—The dry plate is now covered with silver bromide, parts of which have been affected by the light and these can be changed into metallic silver very easily by using some chemical compound to carry out this action.

The parts of the dry plate that were not acted on so strongly by the light will now be changed to so much metallic silver and the parts which were not affected at

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[1] This is done by laying it on a perfectly level marble slab.

all will not be changed at all but will simply remain silver bromide.

This process of photography was discovered by a chemist named Talbot in 1837 when his cat upset a chemical solution of this kind on some papers that he had covered with silver chloride and put in the sunlight to observe the change that the light produced in them. Much to his astonishment he found a photograph upon them of the ceiling.

The chemical solution used to change the affected parts of the silver bromide to metallic silver is known as *hydroquinone* $(KO)_2CrH_4$.[1] When the plate is well soaked in this solution the parts of the silver bromine which were affected by the light from the object are changed into metallic silver and parts which were not affected remain the same.

At the end of the developing process the plate is covered in parts with metallic silver and in others with the silver bromide, which since it was not affected by the light was not attacked by the developer. The metallic silver is black while the silver bromide is white and this makes the picture stand out on the plate so that you can see it.

WHY A FIXING BATH IS NEEDED.—However, if this same plate was again exposed to the light the silver bromide would again be affected and so the picture would be spoiled. For this reason the developing is done in a dark room under the light of a red lamp which will not affect the plate (see Fig. 65). As soon as the plate is developed it is put into a solution which will wash off the silver bromide and so prevent it from being again acted on by the light.

For this purpose a solution of *hyposulphite of soda*, chemically called *sodium hyposulphite* $(Na_2S_2O_4)$, and by photographers simply *hypo*, is used. This solution is known as the *fixing bath*. When the silver bromide on

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[1] Formerly *pyrogallic acid* was used as a developer.

the plate has been washed off it is removed from the bath and the picture then stands out with the white part of the original object black and the black parts of the object white. For this reason it is called a *negative*, for the true relations of light and shade are just reversed.

About Printing from Negatives.—Now since the picture on a glass plate negative has the opposite light and shade values of the object from which it was made it is necessary to make another photograph, or *positive*, as it is called, from the negative, to see the object as it really is.

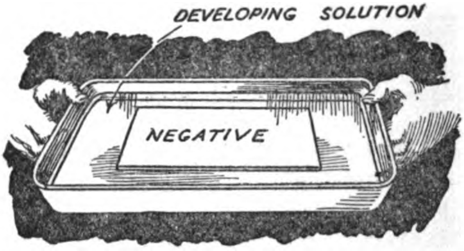

FIG. 65.—DEVELOPING A PLATE IN A DARK ROOM

When a glass positive is made from a negative it is called a *transparency*, but the usual and cheapest way of making a positive is to print it on paper. There are several kinds of photographic printing papers and some print slowly in the sunlight while others print quickly in gaslight but all of them are prepared in about the same way, and very much as dry plates are, that is the paper is coated with silver. The silvered paper is laid on the negative and both are held firmly together in a *printing frame* (see Fig. 66). The frame is then turned toward the sun so that the light shines through the negative and onto the paper. Since the negative is coated with silver in some places and has nothing on it in other places the light passes through the places where there is no coating and strikes the silver bromide on the paper.

The light which passes through the negative turns the paper black and where the light cannot pass through the black parts of the negative which are coated with metallic silver, the paper under them, consequently, is left white. In this way the paper is the opposite of the negative and looks like the original object that was photographed in so far as form and shading are concerned.

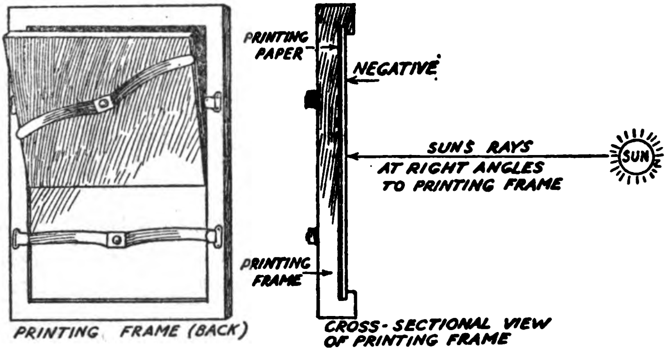

Fig. 66.—How a Printing Frame is Made and Used

The paper now has to be *fixed* in just the same way as a negative and for the same reason, that is to keep it from fading when it is exposed to the light. After being taken from the fixing bath the print is washed in water and dried when it becomes the finished *photograph*. From this you can see that from the time the plate, or film, was made which you slipped with so little thought into your camera, until the moment when you hold the finished picture in your hands great things have been doing and chemistry has been busily at work in aiding you to get a photograph of the object.

Why Films are Used.—In explaining how photography is done I mentioned only the dry-plate process but the same process that is used in making glass dry plates is also used in making *celluloid films* for *kodaks* and moving-picture cameras.

Celluloid films are used because they do not break like the glass plates, they are many, many times lighter and finally they can be rolled up and *loaded* into the camera in broad daylight, whereas glass plates must always be put into the plate holders in a dark room, one at a time.

On the other hand films have the disadvantage of being highly inflammable and sometimes even explosive if they catch on fire in a confined place for their composition is very like that of gun cotton.[1] As the films are small this, fortunately, seldom happens.

Kinds of Printing Papers.—*Velox Paper.*—Silver bromide paper, which is known by the trade name of *velox*, is made by coating paper with an emulsion of silver bromide exactly like dry plates and films. These papers can be printed quickly by an artificial light of any kind and must be developed like a dry plate to bring out the image.

Silver Papers.—Papers that go under the name of *silver* and *solio* are chloride of silver papers; the first is coated with albumen and the second with gelatine and then with the silver chloride. These papers must be printed in the sun and even then it takes from ten minutes to an hour, according to the *density* of the negative, to make a print.

Why Silver Prints are Toned.—To give silver prints a pleasing color they have to be *toned* and some beautiful effects are obtained by this process. The prints are first soaked in *sodium chloraurate*, which dissolves some

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[1] See Chapter on *The Chemistry of Warfare.*

of the silver, displacing *gold* from the chloraurate, which is then deposited on the paper in place of the silver.

The thin film of gold gives a richer color to the print and it has been widely used by photographers. Sodium chloraurate is not the only substance used for toning for *potassium chloraplatinite* acts much in the same manner except that the print is toned with *platinum* instead of gold and this makes a picture that will not fade.

How to Make Blue Prints.—*Blue prints*, as they are called, are used extensively by mechanical engineers and architects because (1) it is a very simple photographic printing process for making duplicate copies of plans, (2) it is quite cheap and (3) the white lines stand out boldly against the blue.

To make blue-print paper get a good quality of drawing paper and coat it[1] first with a solution of *ferric oxalate* and then let it dry. Now place an *India-ink drawing*, a tracing or any other transparent picture over the prepared blue paper, put it in a printing frame and expose it to the strong sunlight.

The ferric oxalate is then reduced by the action of the light to *ferrous oxalate* except where the India-ink lines on the tracing paper protect it, when it still remains in the *ferric state*. When the paper has been printed it is placed in a developer of *potassium ferricyanide* solution.

The ferric oxalate then turns brown and this can be washed out with water leaving the black lines of the tracing white; the ferrous oxalate which turns a deep blue sinks into the pores of the paper and it cannot be washed out and a blue print results.

As a matter of fact the regular blue-print paper which is manufactured and sold by dealers in drawing materials is coated with *ammonium ferric citrate* instead of the oxalate. The *potassium ferricyanide* has also been applied to the paper so that all that is needed to make the print

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[1] This is done by floating the paper on the solution in a tray.

is to expose it with the tracing to the sun, and wash it in clear, cold water. This kind of paper is of course much quicker and better in its action than the experimental paper I told you about above.

HOW TO WRITE ON BLUE PAPER.—Should you want to write on a blue print after it has been made all you have to do is to dip a stick, or a pen into dilute solution of *sodium hydroxide;* this decomposes the *ferricyanide* and when it is washed off it leaves the writing on the paper in white.

CHAPTER XIII

THE CHEMISTRY OF THE SOIL

- How the Soil was Made
- When Life First Came on Earth
- How Plants Live and Grow
- The Fertility of the Soil
- How the Air Acts on the Soil
- What Water Does to the Soil
- Ways to Fertilize the Soil
- The Use of Natural Manures
- About the Rotation of Crops
- What Artificial Fertilizers Are
- Kinds of Artificial Fertilizers
    - Complete Fertilizers, as They are Called
    - Dissolved Phosphates and Potash
    - Wood Ash Fertilizers
    - About Ground Bones
    - And also Dissolved Bones
    - What Tankage is
    - On Dry Ground Fish
    - The Rich Nitrogen Compounds
    - The Value of Blood
    - Cottonseed Meal, Linseed Meal and Castor Pomace
    - The Compounds of Potash
    - Phosphoric Acid Compounds
    - Dissolved Bone Black
    - Thomas Phosphatic Slag
- How Fertilizers are Tested
    - How to Get Samples to Test

CHAPTER XIII

THE CHEMISTRY OF THE SOIL

The most common thing on the face of the earth next to the air which surrounds it and the water which covers three-quarters of it, is the *soil.* In some places the soil is only a few inches deep, while in others it is many feet deep, and as it is from the soil that we get our living let us see how it was made and of what it was made.

How the Soil Was Made.—When our good old earth was thrown off from the sun and had cooled down a little a crust was formed on it of solid rock, and that was several millions of years before man had made his appearance and decided to make his happy home here.

Then as the air and water, and heat and cold acted and reacted on the rocks they were broken up and crumbled to bits and this rock-earth was carried by the streams of water and strewn along in the valleys; those giant rivers of ice, the *glaciers,* were also soil makers for in their slow onward movements and by their tremendous weight they ground and pulverized the rocks and carried and spread them along as they went.

When Life First Came on Earth.—Rocks are largely formed of *silicon* and *oxygen* and when such rocks are pulverized they make up what we call *sand.* But other rocks contain potassium, magnesium, and other metals and nearly all the other elements except *carbon* in such form that it can be used to sustain plant life.

Now since living matter in any form cannot exist without carbon we must conclude that the germs of life which were the first to appear on earth originated in the air and, warmed by the sun, they thrived upon the carbon dioxide that is in it.

Once started these minute organisms lived and died by the billions and as they decayed they were deposited in the soil; together with the moisture and the salts already in the soil better compounds were formed for the production of higher kinds of life such as plants and when these decayed in turn they further enriched the soil. This process has been going on ever since until now much of the ground is covered with a layer of loam as shown in Fig. 67.

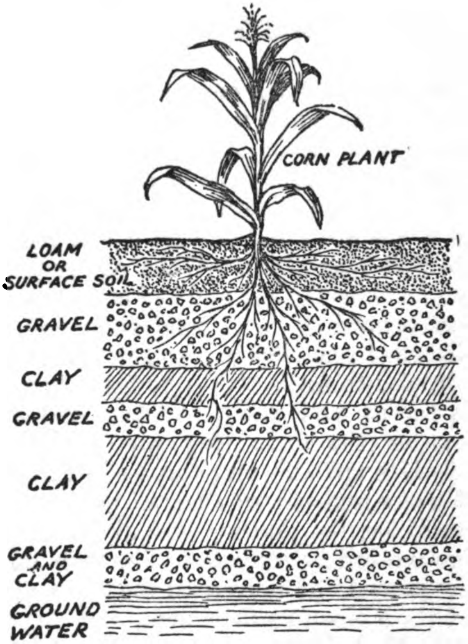

Fig. 67.—How the Soil is Formed

How Plants Live and Grow.—Next let us find out just how plants live and what there is in the soil that they feed upon. To do this it is necessary to know something about the structure of the plants themselves.

In the first place a plant is made up of a large number of little tubes called *capillaries* (pronounced *cap'-il-laries*) and these run from the bottom of the roots up through the stem and out to the tips of the leaves. It is through these important little tubes that the plant must draw its food from the soil, and since it cannot draw solids through them the food must be in liquid form. Cross-sections of a stem are shown at *A* and *B* in Fig. 68.

If you were to examine the walls of the cells which make up these little tubes and the framework of the plant you would discover that they contained granules of starch, and *proteins*, and in the fruit of the plant you would find sugar in one form or another. Since the plant's food is drawn from the soil it is easy to believe that the elements needed to make these foods also come from the soil.

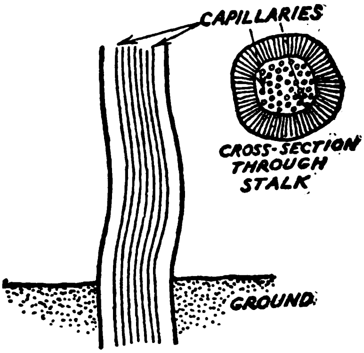

FIG. 68.—THE CAPILLARIES OF A PLANT

Foods That Are in the Soil.—Now starch and sugar contain carbon, hydrogen and oxygen, while proteins contain sulphur, nitrogen and phosphorus. Hence these are the substances which must be in the soil in some soluble form so that they can be dissolved by the water in it and sucked up through the capillaries of the plant to give it life.

Besides these compounds large quantities of carbon dioxide are also necessary since this compound furnishes the carbon that the plant requires.

The carbon dioxide, however, does not come from the soil but the plant breathes it in from the air through tiny openings called *stomatae,* which are located on the under sides of the leaves of the plant as shown in Fig. 69.

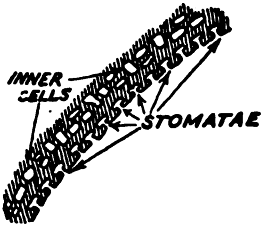

Fig. 69. — Showing the Stomatae of a Leaf Much Enlarged

Finally, as to the other substances I mentioned which are necessary for plant life, since the soil is composed mainly of crushed rock which has come from different localities, and of different plant and animal growths, it contains nearly all of the elements that growing plants must have in one form or another.

For instance, the bones of dead animals contain large amounts of phosphorus, while their flesh, which is chiefly protein, contains sulphur. When they decay both of these substances go back into the soil and there become foods for plants. Verily in this way the lily may in the end feed upon the lion.

Further, any soil has enough magnesium, calcium and iron in it in the form of *bicarbonates* to be able to support plant life. But of all the substances necessary to plant life, nitrogen is the most important, since it must be used to build up the proteins. In plants the proteins are mostly found in the fruits, or those parts of the plants that are eaten, hence their value as food stuffs. You can get a pretty good idea of the value of nitrogen when I tell you that corn requires about 1½ pounds of nitrogen in the ground to grow a bushel of it, and that wheat needs about 2 pounds of nitrogen in the ground per bushel.

The Fertility of the Soil.—When the Pilgrims came to America they found virgin soil for very little of the land had ever been tilled by the Indians and hence this soil contained all the substances needed to make it fertile.

But even at this early date the red men showed the white men how to raise more and better corn by planting a dead fish in each hill with the grains of corn. When the same kind of grain, though, is planted season after season the soil loses its fertility and it becomes necessary for man to either find some way to win back its power of production or else push on and cultivate new soils.

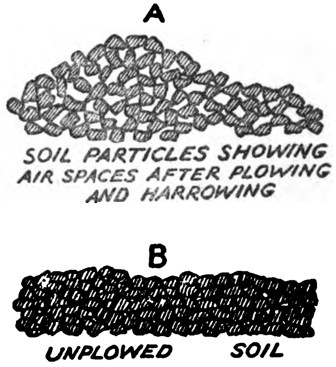

Fig. 70.—What Plowing and Harrowing do to the Soil

How the Air Acts on the Soil.—One way to make soil, good or bad, more productive is to plow it deep and break it up fine so that the air will circulate freely through it as shown in Fig. 70. The way the oxygen acts on the soil is by the oxygen combining with decaying animal and vegetable matter that is in it and these give off phosphoric acid and nitrogen, while the carbon dioxide in the air, which is so necessary to plant growth, is dissolved in the water that is in the soil.

What Water Does to the Soil.—After the water has dissolved the gases, the salts and the acids that are in the soil and has put these plant foods into liquid form, it is sucked up into the plant through its roots.

This it does by a process called *capillarity* (pronounced *cap'-il-lar-ity*) that is the plants absorb it in the same way that a sponge, or a lump of sugar sucks up water.

Moreover the water provides the plant with drink as well as food.

Ways to Fertilize the Soil.—There are four ways by which the soil can be fertilized and these are (1) by *animal manure* (2) by *decaying vegetable matter*, (3) by the *rotation of crops* and (4) by *artificial manures*, or *commercial fertilizers* as they are called.

The Use of Natural Manures.—Natural, or *animal manures* contain large amounts of nitrogen compounds and they are consequently valuable as fertilizers. For this reason raising livestock should always be a part of farming. Barnyard manure is not only valuable as a fertilizer but it keeps the particles of soil apart and this gives the air a better chance to pass through it.

A natural manure formerly much used is *guano*, as the excrement of sea-birds which is found in the dry climate of the coast of Peru is called. Guano is rich in phosphoric acid, nitrogen and potash. *Humus* is a brown or a black mold that is formed by vegetable matter when it decays and this contains acids that go to make for soil fertility.

About the Rotation of Crops.—Some genius finally discovered what might be called a *self-fertilizing* process, and this is known as the *rotation of crops*.

This consists of planting fertile land one season with wheat or corn, or other crop that takes a lot of nitrogen from the soil, and the next season instead of planting wheat or corn again, cow-peas or clover is planted because these plants do not take nitrogen from the soil for their growth, but instead they give it back.

This is done by means of *germs* or *bacteria* as they are called; these little organisms form tubercles (pronounced (*tu'-ber-cls*), that is little swellings on the roots of the plants (see Fig. 71) like cow-peas and clover; they absorb nitrogen from the air and give it to the soil; in this way the soil is greatly improved and the next season the field is again ready to support a crop of wheat or corn.

What Artificial Fertilizers Are.—Sometimes it is not possible to use natural fertilizers, and the soil may be so poor that even the rotation of crops would not make it rich enough. Where such untoward conditions obtain artificial fertilizers must be used. These are made of certain substances that are rich in nitrogen and other plant foods. Since these substances, like those which the earth should contain, must be sucked up through the capillaries of the plant it is clear that they must be soluble, that is easily dissolved.

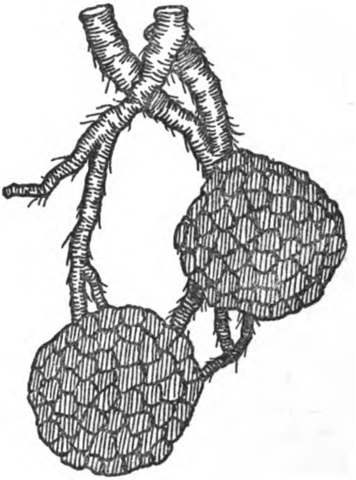

Fig. 71.—Tubercles Growing on Plant Roots

Kinds of Artificial Fertilizers. — So that you can get an insight into this all important branch of farming, I will give you a short description of (1) the various kinds of fertilizers used, and tell you (2) of what they are made.

There are a large number of fertilizers on the market and these can be divided into (1) *complete fertilizers;* (2) *dissolved potash and phosphates;* (3) *wood ashes;* (4) *ground bones, or dissolved bones, tankage and dry ground fish;* (5) *nitrate of soda, dried blood, cottonseed meal, linseed meal and castor pomace:*

Complete Fertilizers, as They are Called.—The fertilizers that come under this head are those that contain the three chief elements which are good plant foods, namely, *nitrogen, phosphoric acid* and *potash.* It is claimed for *complete* fertilizers that they supply the soil with enough food of every kind for plants of every variety. As a rule, though, they fall short in one or more constituents, so don't *bank* too largely on them.

DISSOLVED PHOSPHATES AND POTASH.—These fertilizers are usually made up of *acid phosphate,* formed by dissolving bone black with some potash, the rest of the bulk being made up by adding cottonseed, or some other filler to it.

WOOD ASH FERTILIZER. *Wood ashes* are of course the ashes left after wood is burned. Ashes of this kind supply potash and phosphoric acid to the soil. Of the fertilizers to come all are mixtures containing potash and phosphoric acid which are meant to be mixed with cottonseed meal and other material of a like nature and these are especially good for growing tobacco.

ABOUT GROUND BONES.—Bones contain some nitrogen and a large amount of phosphoric acid. There are two kinds of ground bones used as fertilizers and these are (1) *raw bones,* and (2) *steamed bones.* The raw bones contain more grease than the cooked, or steamed bones do, for the latter have had most of the grease removed by steaming under pressure. As a plant food, however, one is about as good as the other.

AND ALSO DISSOLVED BONES.—To make dissolved bones sulphuric acid is added to finely ground bones, or bone dust, and this makes the bones give up the phosphoric acid more easily. The phosphoric acid in the bones is present in the form of *tri-calcium phosphate;* when treated with sulphuric acid, the tri-calcium phosphate is reduced to two lower forms of phosphate, from which the acid is much easier obtained than from the untreated bones. Dissolved bones of good quality should contain about 2 per cent of nitrogen and between 12 and 16 per cent of phosphoric acid.

WHAT TANKAGE[1] IS.—Like ground bones this material is a refuse product from slaughter houses, and tankage, like bones, is of use because it contains nitrogen and phosphoric acid.

There are many different grades of *tankage* and these are often mixed with other fertilizing agents. Of course the nitrogen contained in tankage is the important element and so the more nitrogen tankage contains the higher value it is to the farmer and hence the more it will cost.

Different grades of tankage are known by numbers which represent the percentage of nitrogen and phosphoric acid in them as follows: *11 and 15 tankage* means that this grade contains 11 per cent of ammonia (that is about 9 per cent of nitrogen), and 15 per cent of bone phosphate (that is about 7 per cent of phosphoric acid). Below is a table showing the grades and the nitrogen and phosphoric acid they contain:

TABLE OF VALUES

Grade of Tankage.	Nitrogen in Tankage, Per Cent.	Phosphoric Acid in Tankage, Per Cent.
11 and 15	9	6–7
9 and 20	7	9
7 and 30	5	14
6 and 35	5	16

ON DRY GROUND FISH.—Wherever fish oil is manufactured the dry ground bodies of the fish from which the oil has been extracted were formerly a waste product. They are now used, however, as a fertilizer because of the unusually high per cent of easily obtained nitrogen they have in them.

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[1] *Tankage* is the name of a kind of fleshmeal made from refuse meat, entrails and other slaughter-house offal.

They also contain considerable phosphoric acid, and there is usually about 8 per cent each of these necessary substances in them.

THE RICH NITROGEN COMPOUNDS.—And now we come to the nitrogen compounds and these find great favor with farmers as fertilizers. The first of these is:

Nitrate of Soda.—This is a compound found in Chile and Peru in large quantities in a crude form and contains anywhere from 2 to 5 per cent of nitrogen. After it is refined it forms light-colored crystals and these contain about 15 per cent of nitrogen.

Nitrate of soda or *sodium nitrate,* to give it its chemical name, is very soluble in water, and it makes an excellent food for plants. Hence, it is used as a *top dressing* for lawns. In order to spread it evenly it is mixed with twice its bulk of earth or some other adulterant.

THE VALUE OF BLOOD.—*Blood* also used to be a waste product of the slaughter houses but as it contains a high per cent of nitrogen it makes a mighty good fertilizer. If the blood is of good quality it will sometimes contain as high as 10 per cent of nitrogen and as it is easily dissolved the plants have no trouble in absorbing it.

COTTONSEED MEAL, LINSEED MEAL, AND CASTOR POMACE.[1]—There are three substances which were formerly waste products but are now used for fertilizers, and they are the meals and pomace left after the cottonseed, linseed and castor beans have had their oils pressed from them.

All of them contain a fairly high percentage of nitrogen. They are largely adulterated with the hulls and this greatly lessens their value for fertilizing for it lowers the amount of nitrogen which they should have for their weight. These substances also contain a little potash and some phosphoric acid:

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[1] The cake which is left over after the oil has been pressed out of castor beans is called *pomace.*

TABLE OF VALUES

	Nitrogen, Per Cent.	Potash, Per Cent.	Phosphoric Acid, Per Cent.
Cottonseed meal..........	7	1	1
Linseed meal..............	6	1	1
Castor pomace............	5	1	2

THE COMPOUNDS OF POTASH.—In past years our potash supply came chiefly from the salt deposits at Stassfurt, Germany, but now we make it ourselves. A good grade of *sulphate of potash* contains about 50 per cent of potash.

In this form it is excellent for the growth of tobacco and also crops which produce starchy products, such as potatoes, sugar, beets, and fruits. *Chlorate of potash* (KCl) also makes a high-grade fertilizer which is well adapted for grass and farm crops in general, while *carbonate of potash*, $K_2(CO_3)$ is fine for tobacco since it contains about 65 per cent of potash.

PHOSPHORIC ACID COMPOUNDS.—When *acid phosphate* is treated with sulphuric acid it makes phosphoric acid (H_3PO_4) and this is an excellent fertilizer. There are two grades of acid phosphate, the first being low and containing only 12 to 15 per cent of phosphoric acid and the second being high and containing from 15 to 19 per cent of phosphoric acid.

DISSOLVED BONE BLACK.—Charred bone powder is used in sugar refineries to take the color out of and to whiten sugar, and when the bone has served this purpose it is made to do duty again as a fertilizer. This is done by treating it with sulphuric acid when phosphoric acid is obtained as explained above.

THOMAS PHOSPHATIC SLAG.— Many iron ores contain phosphorus and the manufacture of free iron from such iron ores leaves what is called *Thomas phosphatic slag*. It is largely used in Germany as a source of phosphoric acid.

How Fertilizers are Tested.—There are a large number of fertilizers on the market and among them are good, bad and indifferent kinds. To know that the kind you are buying is not only good but good for the soil you intend to use it on send a sample to the *Department of Agriculture, Washington, D. C.*, and its chemists will test it for you free of charge.

HOW TO GET SAMPLES TO TEST.—Take a tin tube about an inch in diameter, 3 feet long and open at both ends. Push it into the center of a bag of fertilizer clear to the bottom and pull it out. The contents in the tube will be a fair sample of the fertilizer in the bag.

Several samples should be obtained in this way and then mix each sample well on a clean sheet of paper; divide the sample into two parts and put each part into a separate glass jar. Pack them carefully in corrugated pasteboard tubes and sent them by parcels post to Washington.

CHAPTER XIV

THE CHEMISTRY OF WARFARE

- The Invention of Gunpowder
- What Gunpowder is Made of
  - First Comes Potassium Nitrate
  - About the Other Things in Gunpowder
- What Takes Place when Gunpowder is Fired
- The Reason Explosives Explode
- The Cause of Fouling from Gunpowder
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- Kinds of Smokeless Powder
- What Smokeless Powders are Made of
  - How Nitroglycerine is Made
  - How Gun Cotton is Made
    - Nitroglycerine Powders
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- A Comparison of Smokeless Powders
- How Dynamite is Made
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- How TNT is Made
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- The Further Application of Chemistry to Modern Warfare

CHAPTER XIV

THE CHEMISTRY OF WARFARE

It was at least twenty-five centuries after the chance discovery of gunpowder before chemistry had pushed ahead far enough to improve upon it as an explosive. But when chemistry came to the front during the last century, improvements were in order and explosives having many times the power of gunpowder were invented.

The Invention of Gunpowder.—To the Chinese is given the credit of having discovered how to make gunpowder and this was away back 600 years before the Christian era. From China the news of the wonderful stuff which would *blow up* when it was lighted made its way to Arabia and there the first gun was invented to take advantage of its explosive force.

About the time that Christ was fulfilling His mission on earth the secret of making gunpowder found its way into Italy and some time after that into Germany. It was not, however, until the battle of Crécy[1] in 1346 that gunpowder was used to any extent and it was there and then that the first crude cannons were actually used in warfare.

And so it was on through the Middle Ages and up to recent times that gunpowder was the sole and only explosive known. And although it is such a common substance and is still so largely used, I'll wager that you never have had the slightest idea of how it is made, or of what it is made and why it explodes when fired.

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[1] A town in the Somme *department* of France.

What Gunpowder is Made of.—In the first place, gunpowder is made up of four substances and these are (1) *potassium nitrate*, or *Bengal saltpeter* as it is sometimes called, (2) *sulphur*, (3) *charcoal*, and (4) *water*.

To make 100 pounds of gunpowder requires 75 pounds of potassium nitrate, 14 pounds of charcoal, 10 pounds of sulphur and 1 pound of water. The first three named substances are put into a *mortar* and *ground* with a *pestle* until a perfect mixture has been obtained in which the naked eye cannot distinguish any of its original parts.[1]

Now gunpowder is a mechanical mixture and not a chemical compound,[2] therefore it cannot be represented by a formula. The charcoal is, of course, what gives to the mixture its well-known black color; the water is then added and mixed in with the other ingredients until a heavy paste results, when it is allowed to dry into a hard cake and the cake is then crushed and broken up into grains.

For rifles of small *caliber*[3] the finest grains are used, while for large caliber rifles the coarser grains are employed. Of course when I say it is used for rifles, I mean for loading the cartridges which the rifle shoots. And now before we go further and find out why and how gunpowder explodes let us see just what each of its components is, and something about it.

First Comes Potassium Nitrate.—*Potassium nitrate* formerly was obtained from various districts of India and China, but chiefly from India. The name *Bengal*

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[1] Better let the Maxims or the du Ponts do the mixing unless you are tired of life.

[2] After it explodes, of course it is no longer gunpowder.

[3] The *caliber* of a rifle is the *gauge*, that is, the size of its bore and is measured by the hundredths of an inch. The caliber of big guns is measured either in *inches*, English measure, or in *centimeters* of the metric system.

nitrate came from the potassium nitrate diggings at Bengal, India.

Up to 1825, when the Crimean War broke out, the supply of potassium nitrate had always been plentiful, but at this time it ran short and consequently some other means and ways of getting it had to be devised for this was the only natural supply available.

Now in Chili, South America, there are huge beds of *sodium nitrate*, or *Chile saltpeter*, as it is called, which are about 5 feet deep, 2 miles wide and 220 miles long. The natives call it *caliche* (pronounced *ca-le'-shay*).

Since sodium nitrate is a brother to potassium nitrate it was proposed to convert the former into the latter, by a process of simple double decomposition.

Potassium chloride, which is very plentiful, is heated with the *sodium nitrate* in water, when *sodium chloride* and the desired *potassium nitrate* are formed. The action is as follows:

$$\underset{\text{chloride}}{\overset{\text{Potassium}}{KCl}} + \underset{\text{nitrate}}{\overset{\text{Sodium}}{NaNO_3}} \rightarrow \underset{\text{chloride}}{\overset{\text{Sodium}}{NaCl}} + \underset{\text{nitrate}}{\overset{\text{Potassium}}{KNO_3}}.$$

The NaCl is common table salt and it crystallizes out of the liquor that is left and when this liquor is poured off and still farther cooled the KNO_3 which is wanted then crystallizes out.

Now you may ask why is sodium nitrate not used for making gunpowder instead of potassium nitrate.

The reason it cannot be used is because sodium nitrate is *deliquescent*[1] (see Chapter II) and takes on water from the air while potassium nitrate does not. Since gunpowder must be absolutely dry to burn it is evident that sodium nitrate would not serve the purpose.

ABOUT THE OTHER THINGS IN GUNPOWDER.—And now we come to the other ingredients that go to make up gun-

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[1] That is, it will dissolve gradually and become a liquid by absorbing moisture from the air.

powder. Charcoal is simply *carbon*, and well-burnt wood furnishes this product. You have already learned something about our old friend of alchemist days, *sulphur*, in the chapter on Gases. As a further word I might add that it is the chief constituent of sulphur matches and that it is mined in Sicily.

What Takes Place when Gunpowder is Fired.—Now the action which takes place when gunpowder is fired is a very complicated one and therefore I shall not try to explain it in detail here. It is enough to say that half of the product formed is *potassium carbonate* which is a solid and this accounts for the smoke that is formed, for it is a cloud of little particles of this substance which you see.

Along with this solid product many gases are set free and it is these that are the important products in the firing of gunpowder, for it is the sudden formation and expansion of these gases that forces the bullet from the cartridge shell, shoots it down the barrel of the gun and, finally, makes it travel through the air with a speed of from 2000 to 3000 feet per second. These various stages are shown in Fig. 72.

The Reason Explosives Explode.—Next let us look into the *theory* of explosives just a little and find out why they explode with such great force. All explosives contain nitrogen in combination with various substances in some form or other.

Now nearly all *inorganic*[1] nitrogen compounds, but especially those that are *organic*[2] (that is those which contain carbon) are very *unstable*, which means that they decompose, or break down, very easily. Thus with many explosives such as *dynamite*, it is not necessary to heat

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[1] *Inorganic* compounds are those which do not contain carbon as an essential ingredient.

[2] Oppositely *organic* compounds are those which contain carbon as an essential ingredient and this includes all forms of life.

them in order to start this sudden decomposition, but merely to strike them a sharp blow.

The decomposition, or breaking up, is then so rapid that a tremendous bursting force is exerted by the sudden expansion of the gases formed, and this is the power that drives forth the bullet or projectile.

The Cause of Fouling from Gunpowder.—You will remember that I said that over half of the products of the explosion of gunpowder were solids and that this is what caused the smoke when a gun or cannon is fired. But these solid products have been found to be injurious in several ways to the firearm in which they are formed.

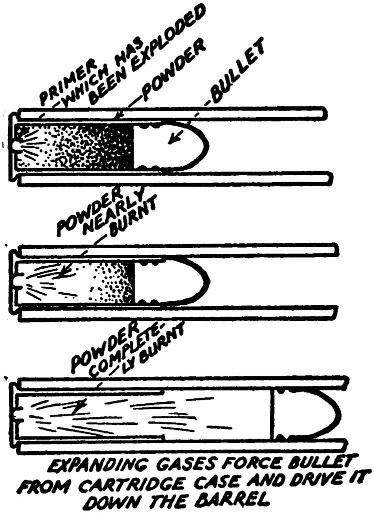

FIG. 72.—HOW A CARTRIDGE IS EXPLODED

The first is that in all guns they *foul* the barrel, that is they stick to the inside of it and as time goes on this coating grows thicker and harder until finally the whole length of the barrel on the inside is covered with a hard, crust or cake.

Since the bullet, or projectile, is of the same diameter as the barrel when clean, it has to force its way through this crust and in doing so it not only loses its *momentum*[1] which is bad, but what is still worse if the crust is thick enough and has become hard enough, particles of lead are scraped from the bullet every time the gun is fired and these soon form a barrier in the barrel which greatly retards the speed of the bullet.

If the *piece*[2] is fired long enough when in this condition the bullet will finally be unable to force its way through the crust and as a consequence all of the expansive force of the gases will be exerted on the barrel near the breech end and the barrel will burst.

Further when *black powder*, that is common powder, was used formerly for cannon and other large guns the great dense clouds of smoke were a *dead give away* of the position of the guns and often furnished valuable information to the enemy.

Since gun positions must be concealed from the enemy's view at all times, and because of the tremendous amount of fouling left by black powder, *smokeless powder* was invented and it was then decided to do away with the former altogether in military practise.

The Value of Smokeless Powder.—When smokeless powder explodes it forms but very little solid matter, most of the products of combustion being gases which are driven out of the barrel. It is easy to see that as a powder for sporting guns and military operations it is far ahead of black powder in every way, with but one exception and this is the tremendous force with which it explodes.

This is due chiefly to (1) the fact that the smokeless

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[1] *Momentum* is the force stored up in a body which tends to keep it going and makes it hard to stop.

[2] A gun is often called a *piece*.

powder does not *burn* as gunpowder does, but explodes all at once, or *detonates* as it is called, and (2) that nearly the whole of the smokeless powder is changed into powerful gases, whereas with gunpowder less than half is changed into gases.

So in the early days of smokeless powder many accidents occurred from this cause although after much experimenting it was discovered how to make it quite safe.

Kinds of Smokeless Powders.—As used to-day for sporting and military rifles there are two kinds of smokeless powders, and these are known as (1) *bulk-for-bulk* powder and (2) *dense* powder.

Bulk-for-bulk powder, as its name shows, is a smokeless powder of which a given bulk is equal to that of a charge of black powder. *Dense* powder on the other hand is so much more powerful than an equal bulk of black powder, that much less of it by bulk is needed for loading. Because these two powders have often been confused there were many cases where cartridges were overcharged to the consequent and complete ruin of those who made the mistakes.

What Smokeless Powders are Made of.—Both bulk-for-bulk and dense smokeless powders are made of two kinds of materials and these are (1) *nitroglycerine* ($C_3H_5(NO_3)_3$) and (2) *nitrocellulose* ($C_{12}H_{14}O_4(ONO_2)_6$).

How Nitroglycerine is Made.—This is the earliest known of the *high explosives.* Unlike gunpowder *nitroglycerine* is a chemical compound and not a simple mixture of substances.

Nitroglycerine is made by treating glycerine with nitric acid when water and *glyceryl nitrate* ($C_3H_5(NO_3)_3$), which is the right name of nitroglycerine, are formed. The formula for this reaction is:

$$\overset{\text{Glycerine}}{C_3H_5(OH)_3} + \overset{\substack{\text{Nitric}\\\text{acid}}}{HNO_3} \rightarrow \overset{\text{Water}}{3H_2O} + \overset{\substack{\text{Glyceryl}\\\text{nitrate}}}{C_3H_5(NO_3)_3}$$

How Guncotton is Made.—When cotton, which is *cellulose* ($C_6H_{10}O_5$), is soaked in a mixture of nitric and sulphuric acids and is washed with water and dried it becomes very explosive; it is then called *guncotton*, or nitrocellulose. This compound does not give off smoke when it burns and it leaves very little ash.

Nitroglycerine Powders.—These powders are made of guncotton which contains between 30 and 50 per cent of nitroglycerine and are very powerful.

Nitrocellulose Powders.—Powders of this kind are practically pure guncotton, and are not as powerful as nitroglycerine powders.

A Comparison of Smokeless Powders.—Nitroglycerine powders burn at a considerably higher temperature than the nitrocellulose powders. The high temperature set up by the former causes the rifle barrel to expand and this opens the pores of it; when this happens the acids set free by the explosion is forced into the pores of the steel and this causes *erosion*, that is they eat it away.

On the other hand nitrocellulose powders burn with much less heat than nitroglycerine powders with the result that the erosion of the steel is not so great and the life of the gunbarrel is considerably prolonged. Finally nitroglycerine powders do not last well in warm climates whereas nitrocellulose powders can be stored in any kind of a climate for any length of time without spoiling.

How Dynamite is Made.—There are many kinds of dynamite but all of them are made by mixing some *inert* material, such as sawdust, or flour, with nitroglycerine when they are formed into sticks.

The purpose of the sawdust, or flour, is to separate the molecules of the explosive one from the other, for the farther apart they are in the mixture the slower will be the explosion and hence the less danger from *detonation*, that is an instantaneous explosion.

How Cordite is Made.—*Cordite* is a kind of smokeless powder. In 100 pounds of cordite there are 65 pounds of guncotton, 30 pounds of nitroglycerine and 5 pounds of *vaseline* and these are dissolved in *acetone*, $(CH_3)_2CO$.

This latter substance is a product obtained by distilling wood and it is a common solvent for all three of the components above named. A paste is formed of the materials and this is rolled out and cut into small pieces. It is then left until the acetone dries out when a horny substance remains and this is *cordite*. Fig. 73 shows several different kinds of powders including cordite.

How Lyddite is Made.—And now we come to two explosives of very like composition and since they detonate, they are largely used for loading bombs, torpedoes and submarine mines; named these are (1) *lyddite* and (2) TNT.

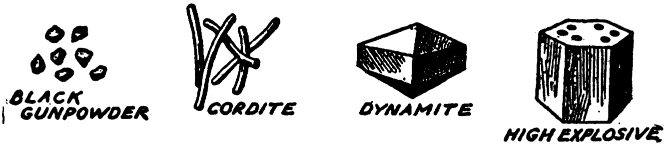

FIG. 73.—FORMS OF POWDERS.

Picric acid, or trinitrophenol as the chemists call it, is formed by heating nitric acid with *phenol*, which is the chemical name of *carbolic acid*, and the *lyddite* crystallizes out of the solution in the form of yellow needles. The crystals are used as a dye as well as to make people die.

How TNT is Made.—This is the explosive that is generally used for charging the *warhead* of a torpedo, as that part of it which explodes when it strikes the doomed ship is called.

Now you have probably often wondered what these mystic letters which you have seen so often in the news-papers mean and here it is. TNT is the abbreviation for a compound known as *trinitrotoluene* and which is made by heating nitric acid with *toluene* ($CH_3C_6H_5$), a reddish organic compound that belongs to the same family as *benzene*, and you know how that blows up, thus:

$$\overset{\text{Toluene}}{CH_3C_6H_5} + \overset{\text{Nitric acid}}{3HNO_3} \rightarrow \overset{\text{Water}}{3H_2O} + \overset{\text{Trinitrotoluene}}{CH_3C_6H_2(NO_2)_3}.$$

TNT is great stuff for filling not only torpedoes but high explosive shells for big guns as well, because shocks will not set it off. The fact that it can be melted without danger when it can be poured into the shells makes filling them easy, safe and rapid in the extreme. For the weapons of modern warfare it makes an ideal explosive and huge quantities of it are manufactured and used.

When TNT Explodes.—When TNT explodes a large amount of black smoke is formed because of the carbon which is set free by the decomposition thus:

$$\overset{\text{Trinitrotoluene}}{2CH_3C_6H_2(NO_2)_3} \rightarrow \overset{\text{Water}}{5H_2O} + \overset{\text{Nitrogen}}{3N_2\uparrow} + \overset{\substack{\text{Carbon}\\\text{monoxide}}}{7CO\uparrow} + \overset{\substack{\text{Free}\\\text{carbon}}}{7C}.$$

Now as I said before TNT cannot be exploded by concussion as guncotton and nitroglycerine can but it must be fired by means of what is known as a *detonator*.

Since detonators are used on all kinds of explosive weapons from the .22 caliber cartridge rifle up to the new *boche* gun that shoots 74½ miles it is both meet and fitting that you should know something of their action in order to make your knowledge of explosives a little fuller.

What Detonators are and How they Work.—A detonating compound consists of certain chemicals which decompose rapidly and with great force when it is struck a sharp blow.

This compound is molded into *caps*, and these are set in a holder known as a *priming cup* and around which the charge of real explosive is placed. A *hammer*, or striker, is operated by a trigger and spring just as in a rifle or a revolver. When the trigger is pulled, or otherwise released, the hammer is driven forward by the spring and strikes the detonating charge in the cap; the charge then *detonates* and this in turn sets off the explosive charge as shown in Fig. 74.

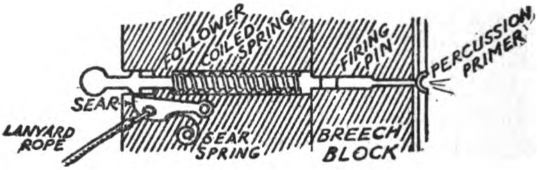

FIG. 74.—HOW A DETONATOR IS OPERATED

On a torpedo this mechanism is known as the *pistol* and it is the striker hitting against the hull of the ship to be torpedoed that explodes the detonator. To prevent the striker from hitting the detonator accidently a *butterfly nut*, that is a nut with wings on it, is fastened to one end of it as shown in Fig. 75.

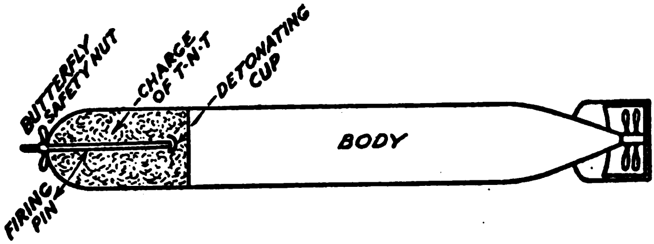

FIG. 75.—THE WAR HEAD OF A TORPEDO

As the torpedo speeds through the water the action of the water on the butterfly nut unscrews it and it finally drops off and leaves the striker free to act.

There is only one kind of detonating compound in common use for exploding cartridges, torpedoes, mines, bombs and shells, and this is *mercuric fulminate* ($Hg(ONC)_2$), or *fulminate of mercury* as the old school chemists called it. It is made by treating mercury with nitric acid and adding alcohol to the solution when a dense white precipitate is formed, which is $Hg(ONC)_2$. This is dried and made into *percussion caps* by the addition of a *binder*, that is gum arabic or some other substance, which will hold it together.

What Poison Gases are Made of.—One of the *Kultured* ideas the *boches* got up to kill off the Allied armies was to poison them to death. In all they have used at least twenty different *poison gases* and every little while they add a new one to the list.

The First Poison Gas.—This was just common *chlorine* and so gas-masks which contained an alkali carbonate was used to offset it. This poison gas was soon followed by *phosgen*, that is *carbonyl chloride*. As a protection against it *sodium phenate* was used in the gas masks.

Tear and Sneezing Gases.—Gases to make the eyes water so that our soldiers could not see were made first of *benzyl bromide*, then of *trichloromethyl chloroformate*, and finally of *phenyl carbylamin chloride*. Then came a *sneezing gas* to make the boys sneeze their gas masks off and this was *diphenyl chloroarsin*.

Mustard Gas.—This gas is the worst one yet invented by the Huns. As many as 50,000 shells filled with this gas which is *dichloro-diethyl-sulphide* have been hurled into the trenches of the Allies in a single night.

At first this gas, which smells like garlic, hurts the eyes of the victim, and then after a few hours they begin to burn, swell and blister. The pain is fearful and to add to the suffering it causes coughing and vomiting.

The Further Application of Chemistry to Modern Warfare.—But the application of chemistry to modern warfare does not by any means stop at explosives and poison gases, for nearly everything that is made or is done has to do with chemistry.

Thus in making steel from which big guns are turned and armor plates are rolled, preserving food for soldiers' rations, manufacturing their clothing and dyeing it olive-drab, or O. D., as it is called, and in a thousand and one other ways, both great and small, chemistry is employed and plays a mighty part.

But after all it is of the most vital importance in manufacturing explosives, as you have seen, for without them there would be no war at least on the gigantic scale that is being carried on now.

CHAPTER XV

SYNTHETIC CHEMISTRY FOR BEGINNERS

WHAT SYNTHETIC CHEMISTRY IS
HOW SYNTHETIC CAMPHOR IS MADE
HOW SYNTHETIC INDIGO IS MADE
HOW SYNTHETIC TANNIN IS MADE
HOW SYNTHETIC RUBBER IS MADE
HOW SYNTHETIC DIAMONDS ARE MADE
HOW SYNTHETIC EMERALDS ARE MADE
HOW SYNTHETIC RUBIES ARE MADE
HOW SYNTHETIC SAPPHIRES ARE MADE

CHAPTER XV

SYNTHETIC CHEMISTRY FOR BEGINNERS

THERE are three ways by which a compound can be produced and these are by (1) the *natural* way; (2) the *artificial* way, and (3) the *synthetic* way.

Take for instance the compound we call *camphor;* when it is obtained from the sap of the camphor tree it is, of course, produced by *nature* and hence it is *natural camphor*. If now you will take various ingredients and mix them together so that the compound has about the same look and smell of natural camphor you will have an *artificial camphor*.

What Synthetic Chemistry is.—But if you know just what elements natural camphor ($C_{10}H_{16}O$) is made up of, and you can take these either separately, or as they are found in something else, and combine them in exactly the same proportions you will have made *real camphor* but it is *synthetic* because *you* built it up instead of nature.

After chemists were able to analyze and find out just what elements each natural substance and compound is formed of it would seem to be quite a simple matter to take the same elements and combine them in the same proportions that nature had and so get the same substance or compound.

But this is far from the truth for there are very few substances made by nature that chemists can duplicate in their laboratories and most of these are long and complicated processes; but nearly everything that chemists do make synthetically they make as good or better than nature herself and they make it cheaper.

For the reasons given above you must not expect to make compounds by *synthesis*, at least until you are a real chemist, but what I shall tell you about this branch of chemistry will start you out right and make you *think*. If this be the case you may make something in due time that will rival nature or perhaps go her one better.

How Synthetic Camphor is Made.—For many years *camphor* ($C_{10}H_{16}O$) has come from the Island of Formosa where it was obtained by chopping down camphor trees. cutting them into bits and distilling them. As the Japanese Government controlled this industry we had to buy the product at a high price or else do without it.

About fifteen years ago, however, an American chemist, named Lovejoy, was struck by the odor of camphor that came from an oil which was a product of turpentine. Now the formula for turpentine is $C_{10}H_{16}$ and from this the chemist reasoned that if he could add an atom of oxygen to turpentine he would have a substance with a formula like camphor and it would be camphor in very truth.

After long experimenting he found a way to add the atom of oxygen to the turpentine and so made *synthetic camphor*. The first step in the process was to put a ton or so of *turpentine* into a steam vat, lined with *asbestos* to keep in the heat, together with some *oxalic acid* which is rich in oxygen. *Pinol oxalate* and *pinol formate* are formed by this action and then a hydroxide is put in and the steam turned on.

The liquid is now distilled and several sweet smelling oils are formed and this leaves behind what is called *borneol* which is a kind of camphor. The borneol is then distilled by a process called *fractional* distillation which means that at different temperatures different oils, or *fractions*, are driven off and these are collected. These oils are used for perfumes and soaps.

The borneol camphor has now changed to real camphor but it is still not the equal of that made by nature. To make it so it is washed by forcing it under pressure through a filter press and this removes all remaining traces of oil in it. The camphor is then put into an oxidizing tank and the oxidizing liquors are thrown off by a centrifugal machine in which the camphor remains behind because it is heavier.

After this process the camphor is still of a light-brown color and to get this out it is slowly evaporated which drives out every trace of water in it. When the camphor has been heated enough a blast of air is forced across the pan containing it when particles of it are blown into a crystallization chamber where they form snow white flake crystals.

This completes the chemical operations in making synthetic camphor and all that remains to be done is to melt the crystals and mold them into cakes. Thus in a few hours man can do that which it takes nature several years to accomplish.

How Synthetic Indigo is Made.—*Indigo* ($C_{16}H_{10}N_2O_2$) as you know is used at home for whitening clothes but what you may not know so well is that more of it is used as a dye than any other kind of a dye-stuff.

It was made in time past of the indigo plant which was largely cultivated in different parts of Europe, in Egypt and then in India but nearly all of it that has been used since 1907 is made in chemical labs.

The starting point for making synthetic indigo is *naphthaline* ($C_{10}H_8$) and this is a coal-tar product. Now don't let the jaw-breaking chemical names get on your nerves but just think what the chemists must have suffered when they invented the process and be glad.

When heated with *concentrated sulphuric acid* and mercury, the naphthaline is oxidized and it then becomes *phthalic* (pronounced *thal'-ic*) anhydride. This latter substance is next treated with an *alkaline hypochlorate* when it is changed into *ortho-ammobenzoic acid.*

By condensing this acid with *monochloracetic acid* an acid is formed that is called *phenylglycocollortho-carboxylic* acid; when this acid is melted with *caustic potash,* it gives *indoxylic* acid and as this acid loses its *carbon dioxide* it changes into *indoxyl* which is getting pretty close to indigo. By treating the indoxyl with an alkaline solution it is changed to *indigotine* which is *indigo,* that is *indigo blue.*

From the above operations it might seem that it would be more costly to make synthetic indigo than to get it by growing indigo plants, but this however is not so for it is, in fact, very much cheaper and it has all but displaced the natural product.

How Synthetic Tannin is Made.—*Tannin* or *tannic acid* ($C_{14}H_{10}O_2$) is a substance used chiefly in the manufacture of leathers, though it is also used for bleaching. The natural product is obtained from the *sumac* bush.

Synthetic tannin is made by heating *phenol,* that is carbolic acid, with *formaldehyde* (CH_2O) in a slightly acid solution and dissolving the resinous products obtained in this way with sulphuric acid. The tannic acid thus obtained is used to bleach dark colored leathers and does not make them lose weight as the natural tannin does.

How Synthetic Rubber is Made.—Many attempts have been made to produce synthetic rubber but none of them have resulted in making a rubber that is equal to the natural product. The synthetic rubber here described is however one of the best.

Rubber, that is *raw rubber* or (C_5H_8) or *caoutchouc* (pronounced *ca-chuc'*) is made by exposing the milky juice of the tropical rubber tree to the air when it thickens. *Elastic rubber* is made by heating raw rubber with a little sulphur and this process was discovered by Charles Goodyear many years ago.

Now there is a substance known as *isoprene* (pronounced *i-so-prene*) (C_5H_8) that has the same formula as raw rubber but there is a difference in the arrangement of their molecules. When isoprene is heated with *sodium*, or some other contact agent, it changes into raw rubber and the latter is then heated with ⅛ of 1 per cent of sulphur.

Isoprene belongs to a series of chemicals known as the *unsaturated hydrocarbons* and among these are *ethylene*, *acetylene*, *benzene* and *toluene*.

How Synthetic Diamonds are Made.—The first chemist to actually make real diamonds was Moissan, a Frenchman. As you may know the diamond consists of pure carbon (C)[1] which has been subjected to great heat and pressure and which has changed it into a hard and brilliant crystalline form that is so much prized in gems.

At various times diamonds have been discovered in fallen meteors which consist chiefly of iron. When the meteors struck the ground they quickly cooled and this made them contract and exert great pressure on the melted carbon inside of them and this crystallized the latter and made diamonds of them.

Moissan concluded that if nature could make diamonds in this way he could do it too. The first thing he had to do was to get a heat great enough to melt the carbon. This he did by using an electric furnace in which an arc-light was formed between two carbon electrodes an inch in diameter and a foot long.

This arc-light used a current of 1000 *amperes* at a pressure of 500 *volts* which was equal to using about 670 horse-power. The arc-light which was enclosed in the furnace developed a heat of between 3000 and 6000 degrees Centigrade. A cross-section view of the furnace

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[1] The diamond, carbon, and graphite are all pure carbon, but in absolute contrast to the diamond, graphite is very soft. Lead pencils are made of graphite.

is shown in Fig. 76, while *A* and *B* in Fig. 77 show the furnace dissected and when set up ready for work.

The walls of the furnace were made of limestone as this substance will resist high temperatures. Now into a crucible the chemist put a ¼ of a pound of pure Swedish soft iron together with the same amount of powdered carbon and these he covered over with powdered charcoal.

The crucible was then set inside the limestone furnace and the carbon electrodes were passed through holes to the inside and directly into the crucible as shown in Figs. 76 and 77 so that the heat from the arc would strike directly on the iron and carbon content.

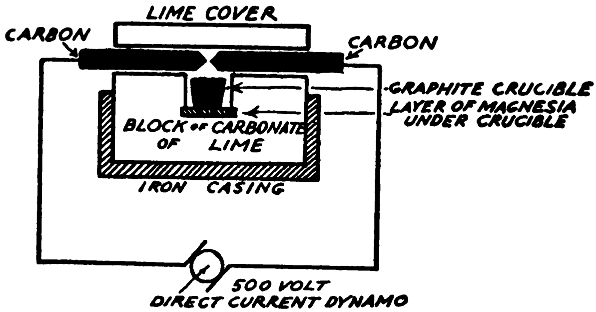

FIG. 76.—CROSS SECTION OF ELECTRIC FURNACE FOR MAKING DIAMONDS

Then the current was turned on and the heat grew greater and ever more intense until the iron and carbon were brought to a melting point and after a little time they formed a clear fluid with little more color than water has.

The next step was to cool the iron so that it would exert sufficient pressure on the carbon to cause it to crystallize into diamonds. This was done by suddenly grasping the crucible with a pair of tongs, taking it from the furnace and plunging it into a tub of cold water. The cooling shrinks the iron so suddenly and the pressure is so great that the melted carbon in it crystallized and diamonds resulted.

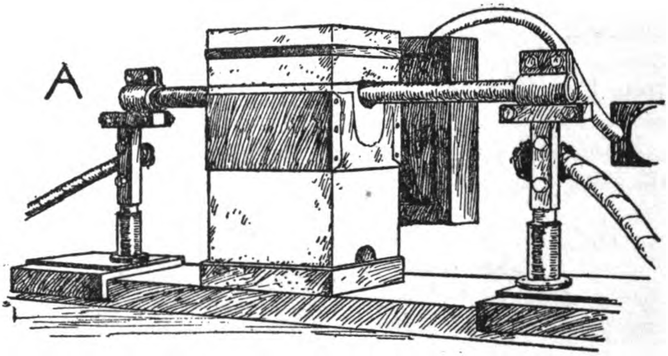

FIG. 77A.—MOISSON ELECTRIC FURNACE FOR MAKING DIAMONDS

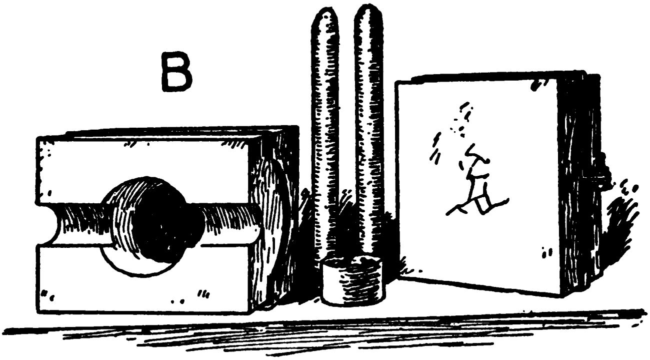

FIG. 77B.—THE SAME DISSECTED

When the iron had entirely cooled off it was dissolved away by hydrochloric acid and this left only the rough carbon that was mixed up with the diamonds. This carbon was dissolved by putting it in a solution of nitric and hydrochloric acid, or *aqua regia* as it is called.

The carbon which still remained was then heated in sulphuric acid when potassium nitrate was forced into it and this took off more of the waste carbon around the diamonds. Finally the last particles of carbon were removed from the diamonds by treating them with *methylene iodide*. The diamonds obtained are shown in Fig. 78.

Making diamonds sounds good but unfortunately they are too small and too costly to be of commercial value and no way has yet been discovered to make larger and cheaper ones.

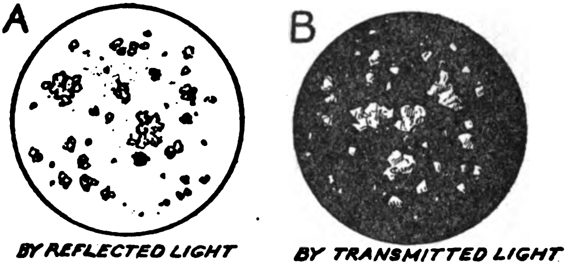

Fig. 78.—Artificial Diamonds Viewed by Reflected and Transmitted Light

How Synthetic Emeralds are Made.—Like synthetic diamonds beautiful synthetic emeralds are made in the laboratory but they are both too small and too costly to make their manufacture profitable.

The *emerald* is a variety of beryl ($Al_2Gl_3(SiO_3)_6$) and when it is clear and flawless is highly prized as a gem. It is known in chemistry as a *metasilicate* of *aluminum* and *glucinum*, or *beryllium* as the latter is called.

About 1890 Haute-Feuille and Perry made some synthetic emeralds and they went about it like this: They dissolved aluminum and glucinum, in the proportions which make up the real emerald, in a bath of *lithium dimolybdate*.

This was kept at a temperature of about 400 degrees Centigrade for fifteen days when little emeralds of perfect composition and appearance crystallized out of the solution.

How Synthetic Rubies are Made.—And now we come to some synthetic gems that are made in the *oxyhydrogen furnace* and which are as clear and large and beautiful as any nature ever produced in her workshop and, glory be, which are quite cheap to make.

The ruby is a gem of pigeon blood-red color and it is the first of the precious stones to be made synthetically on a commercial scale. The ruby is simply *oxide of aluminum* (Al_2O_3), which is a substance only 1 degree less hard than the diamond, and to which has been added a trace of *chromic oxide* (Cr_2O_3) which is a green powder and it is this latter substance, strangely enough, that gives the gem its beautiful blood-red color.

While in the rubies that nature makes traces of other substances are often found the slightest trace of any other substance in a synthetic ruby will completely change the gem and sometimes spoil it entirely. Hence by using other oxides than chromium blue gems, that is *sapphires*, can be made and gems the like of which nature never produced are now made by the chemist synthetically.

The heat used for making synthetic rubies and other gems of this class is obtained by means of an oxyhydrogen flame enclosed in a furnace and temperatures ranging from 1900 degrees to 2400 degrees Centigrade are had.

The blowpipe which makes the flame is formed of two tubes one inside the other, the inner tube carrying the oxygen and the outer tube carrying the hydrogen as shown in Fig. 79, and Fig. 80 shows the furnace complete. The inside tube projects a foot beyond the end of the outside tube and it has a cylinder for a top. In the lower part of the torch the hydrogen tube extends an inch or two beyond the oxygen tube.

The mixture of aluminum and chromium oxides is put into a box which has a sieve on the bottom of it and this

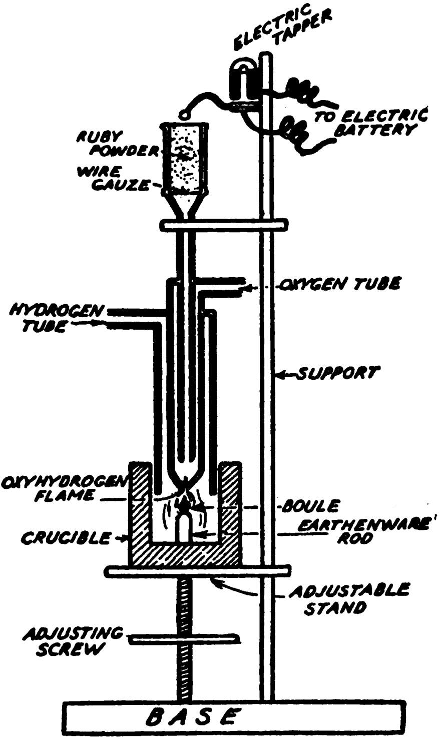

FIG. 79.—THE OXY-HYDROGEN FURNACE

is screwed to the top of the oxygen tube of the blowpipe torch. The torch is then lighted, and a little electrical hammer is set going which taps the top of the box and this causes the particles of aluminum to fall through the flame and onto an earthenware rod which sets directly under the box so that it will catch the heated particles.

When the operation is started the flame heats the rod and as the powder falls it builds up a little pyramid on top of it. As the flame gets hotter the tip of the pyramid begins to melt and a little ball is gradually formed on it.

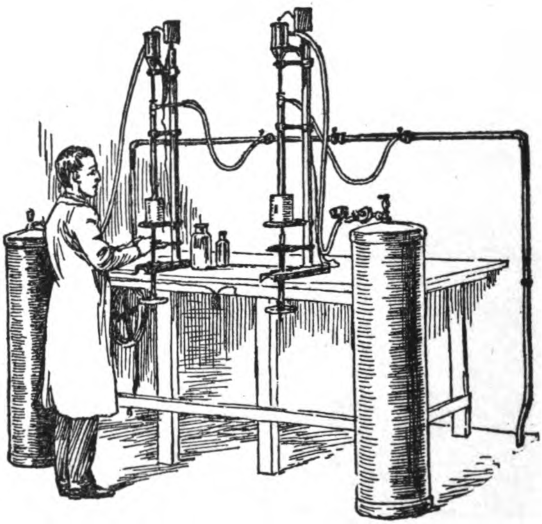

FIG. 80.—MAKING A RUBY IN AN OXY-HYDROGEN FURNACE

The flame is given more oxygen and is made still hotter and as the powdered oxides fall through the flame they melt and collect in drops on the little ball, or *globule* as it is called. The bottom of the globule grows larger and larger until finally a pear-shaped mass called a *boule* (pronounced *boo-la'*) is formed as shown at *A* and *B* in Fig. 81.

This boule is a single crystal and when it is cut it is just like any ordinary gem and if the ruby making mixture of aluminum was pure a clear pigeon-blood ruby will be produced. The cut gems then look like those shown at *C*.

The only way to tell a synthetic gem from a real, genuine natural gem is by the flaws that are in it, that is to say in a synthetic gem the sides of the flaws are round for the flaws are really only bubbles of air, while in a natural gem the flaws have flat sides. However, it takes a magnifying glass of considerable power to detect this difference.

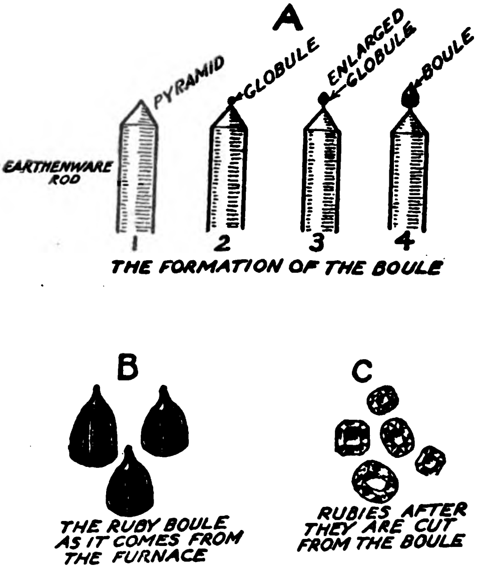

FIG. 81.—PROCESS OF FORMING SYNTHETIC RUBIES

How Synthetic Sapphires are Made.—Sapphires are made of aluminum oxide just as rubies are but they are given their blue color by mixing in a little of the *oxides of tin* (SnO and SnO_2) and *titanium oxide* (TiO_2). The reason all of these oxides are used is to make them keep the blue color which otherwise has a tendency to disappear.

The process of making sapphires is exactly the same as that described for making rubies, that is an oxyhydrogen flame is used and the gems are formed in the same way. When a synthetic sapphire is cut it is next to impossible to tell it from one made when the earth was young by mother nature.

Every year there are 10,000,000 *carats*[1] of synthetic rubies and 6,000,000 carats of synthetic sapphires made and the demand is constantly growing.

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[1] A *carat* is the unit used for weighing precious stones. The carat in general use is equal to 3.168 grains, or 205 milligrams.

APPENDICES

APPENDIX I

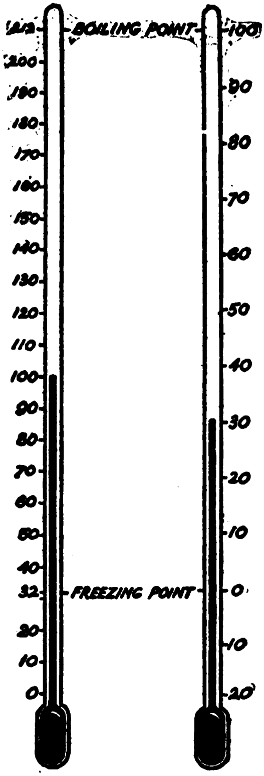

FIG. 82.—FAHRENHEIT AND CENTIGRADE THERMOMETER SCALES COMPARED

APPENDIX II

HOW ACIDS AND THE SALTS THEY FORM ARE NAMED

The *ic* Acids.	The *ous* Acids.	The *hypo* Acids.	The *hydro* Acids.
Those acids containing the most oxygen in combination as HNO_3, nitr*ic* acid $HCLO_3$, chlor*ic* acid H_2SO_4, sulphur*ic* acid are the *ic* acids.	Those acids containing *one less* atom of oxygen than the *ic* acids as HNO_2, nitr*ous* acid $HClO_2$, chlor*ous* acid H_2SO_3, sulphur*ous* acid are the *ous* acids.	Those acids with still less oxygen in them than the *ic* or *ous* acids, have the prefix *hypo* thus $HClO$, *hypo*chlor*ous* acid.	Acids which have no oxygen in them take the prefix *hydro*, as HCl, *hydro*chlor*ic* acid

THE ABOVE ACIDS FORM THE FOLLOWING SALTS

Salts ending in *ate*	Salts ending in *ite*	Salts Ending in *ide*
as Copper sulph*ate*	as Potassium sulph*ite*	as Zinc chlor*ide*

APPENDIX III

SOLUBILITIES OF CHEMICALS

Potassium, ammonium, and sodium.	Carbonates, *except* sodium, potassium and ammonium.
Acetates, nitrates and chlorates.	Oxides and hydroxides, *except* ammonium, sodium, potassium, barium, and calcium.
Chlorides, *except* silver, mercury, and lead.	
Sulphates, *except* barium, calcium and lead.	

APPENDIX IV

SOME ELEMENTS, THEIR SYMBOLS, ATOMIC WEIGHTS AND VALENCES

Element.	Symbol.	Atomic Weight.	Valence.
Aluminum	Al	27	III
Antimony	Sb	120	III, V
Argon	A	40	
Arsenic	As	75	III, V
Barium	Ba	137	II
Bismuth	Bi	208	III, V
Boron	B	11	III
Bromine	Br	80	I
Cadmium	Cd	112	II
Calcium	Ca	40	II
Carbon	C	12	IV
Chlorine	Cl	35	I
Chromium	Cr	52	II, III, IV
Cobalt	Co	59	II
Copper	Cu	63	I, II
Fluorine	F	19	I
Gold	Au	197	I, III
Helium	He	4	
Hydrogen	H	1	I
Iodine	I	127	I
Iron	Fe	56	II, III
Lead	Pb	207	II, IV
Lithium	Li	7	I
Magnesium	Mg	24	II
Manganese	Mn	55	II, IV
Mercury	Hg	200	I, II
Nickel	Ni	59	II
Nitrogen	N	14	III, V
Oxygen	O	16	II
Phosphorus	P	31	III, V
Platinum	Pt	195	IV
Potassium	K	39	I
Silicon	Si	28	IV
Silver	Ag	108	I
Sodium	Na	23	I
Strontium	Sr	87	II
Sulphur	S	32	II, IV, VI
Tin	Sn	119	II, IV
Zinc	Zn	65	II

INDEX

Pres.
Dean Alfred
Vivian