SIMPLE CHEMICAL EXPERIMENTS

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Illustrated by

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CHAPTER I

YOUR LABORATORY

MOST of the experiments described in this book can be performed without elaborate equipment or apparatus. For them you will need only a few bottles, test-tubes, measuring-spoons, and an alcohol lamp. Jelly glasses, mayonnaise jars, small enameled saucepans, and thin glass tumblers can often be substituted for the beakers, flasks, and glassware of the professional chemist.

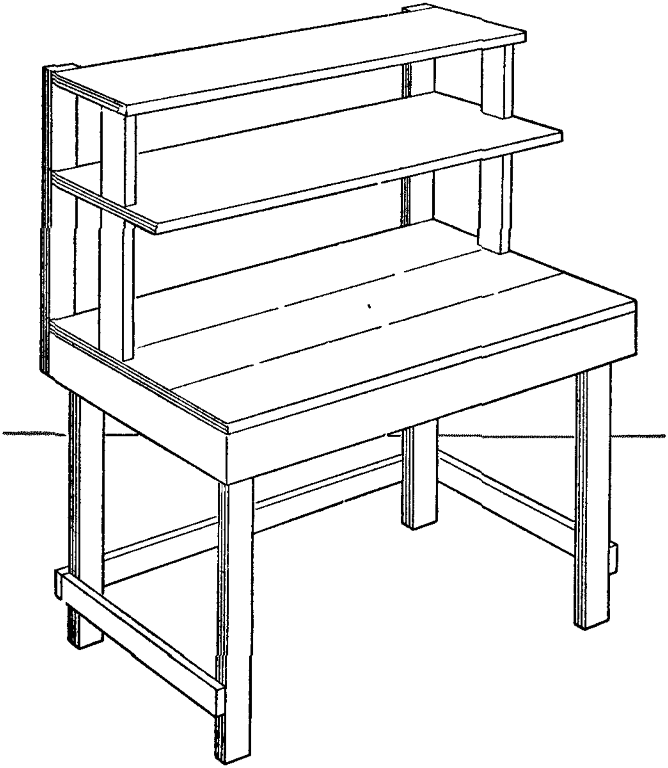
A few of the experiments require beakers, flasks, tubing, funnels, filter paper, crucibles, mortar and pestle, and Bunsen burner. The small sizes of these are not expensive. Frequently the cost of apparatus and chemicals can be shared by establishing a "community" laboratory which is used by two or more experimenters.

Of course it is more convenient to use commercial chemical equipment and more "professional" than to use the home-made variety.

Several firms which advertise in the columns of the popular mechanical and scientific magazines will fill mail orders for glassware and other apparatus. Four or five dollars will equip a home "lab" well.

THE LABORATORY BENCH

For occasional experimenting, the kitchen table makes a good place. But the experimenter who devotes a good deal of time to his hobby or who owns a considerable assortment of chemicals and apparatus needs a table or bench where he can leave his equipment and where it will remain undisturbed.



A design for a simple laboratory bench which can be built from packing-case lumber. A hammer and saw are the only tools required.

A table, a shelf, and a bucket or large can for waste material are the essentials in arranging a home laboratory. When there is a choice of locations, the spot selected should be one where there is no danger to rugs and furniture. If possible, it should be near running water, a gas outlet, and a sink.

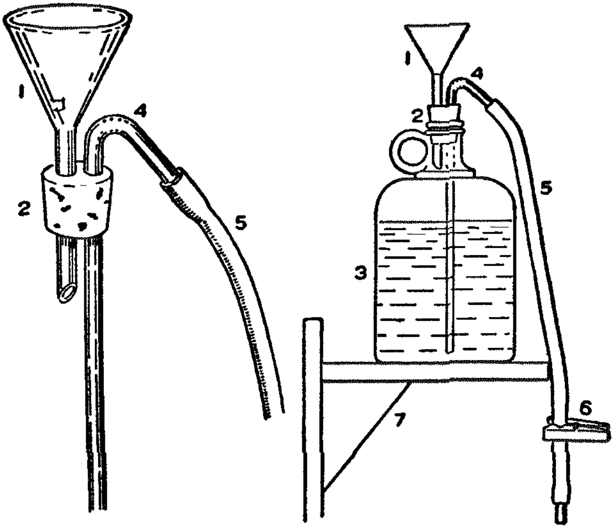
An old kitchen table, a home-made table, even a large packing-case, when fitted with shelves, is a satisfactory laboratory bench. Here much that is interesting about one of the most important and most useful of modern sciences may be learned. A coat of asphaltum paint will protect the surface of the table and make it resistant to water, acid, and other chemicals which may boil over or be spilled during an experiment. Cover that portion of the bench where the Bunsen burner or alcohol lamp is used with sheet-asbestos.

A SIPHON BOTTLE

When running water is not available, a siphon bottle for storing and delivering water at the laboratory bench will be found useful. This can be rigged out of a one-gallon glass jug, a funnel, rubber tubing, a two-hole stopper, glass tubing, and a spring clothes-pin or pinch-cock. The whole arrangement is shown in the illustration. The siphon bottle should be set on a shelf or bracket above the bench level.

To start the siphon, fill the bottle by pouring water in the funnel. Then tilt the bottle until the water fills the glass tube and begins to flow out of the rubber tube. Clamp the spring clothes-pin or pinch-cock over the rubber tube so as to shut off the flow of water and set the bottle on its shelf above the level of the "lab" bench. After the siphon has been started, water will flow out of the rubber outlet tube when-

SIPHON BOTTLE

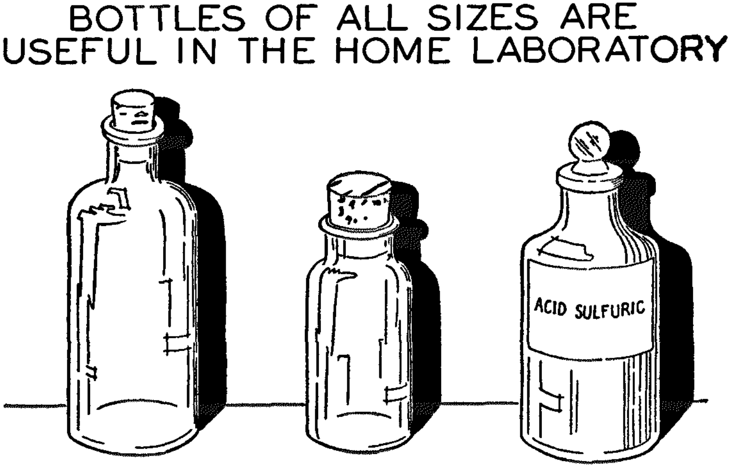


The amateur chemist who does not have running water near his laboratory table should rig up a siphon bottle. A glass funnel (1) and a bent glass tube (4) are inserted in a 2-hole cork (2) which fits into a 1-gallon glass jug (3). The end of the tube inside the jug should reach almost to the bottom. A rubber tube (5) is fitted to the outside end of the glass tube. It should hang down below the bottom of the jug and be fitted with a pinch-cock or spring clothes-pin (6) which will clamp it shut. Fill the jug and tube with water and place it on a shelf (7) at the back of the laboratory table. Water may be drawn from the jug by releasing the clothes-pin or pinch-cock.

ever the pressure of the spring clothes-pin is removed, unless the bottle is empty. The water in the bottle can be replenished by pouring water into the funnel.

BOTTLES

Empty bottles of all sizes and shapes are needed in the home laboratory. Some will be useful for storing and preserving your chemicals. Others can be used as part of your laboratory apparatus. Small, wide-mouthed bottles and small jars can be used for storing powdered or crystallized chemicals, test papers, etc. The ordinary narrow-necked or small-mouthed bottle will serve best for liquids.

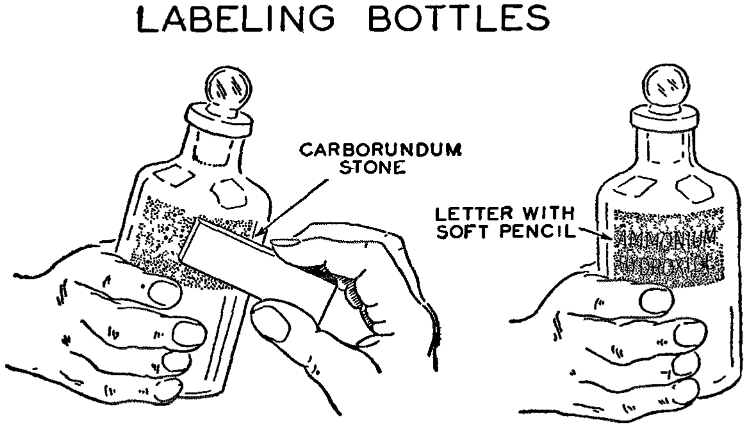


Wide-mouthed bottles make the best containers for powders and crystals. Liquids and solutions can be poured most easily from narrow-necked bottles. Acids and solutions of strong alkalis (sodium hydroxide, ammonium hydroxide, etc.) should be kept in bottles provided with rubber or glass stoppers.

Old bottles must be thoroughly washed, rinsed, and dried before using. A bottle brush will assist greatly in the cleansing process. Bottles in which acids are stored should have glass or rubber stoppers. Sulfuric, nitric, and hydrochloric acids will soon decompose a vegetable cork.

LABELING BOTTLES AND JARS

All bottles and jars containing chemicals should be labeled with the name of the contents and, if you wish, with the chemical formula for same. This is a good way to become familiar with the formulas that chemists use. A formula indicates a single molecule of a compound. It shows at a glance just what the elements of a compound are and how they combine.



This is a simple method of preparing a bottle or other piece of glassware so that you can write on it with a pencil. The surface of the glass is roughened or frosted by rubbing with a Carborundum stone kept wet with water. The pencil marks can be rubbed off with a damp cloth.

A bottle is easily labeled with a strip of paper upon which the name and formula for its contents are written. The label can be held in place and at the same time protected from stains with a piece of transparent "Scotch" cellulose tape.

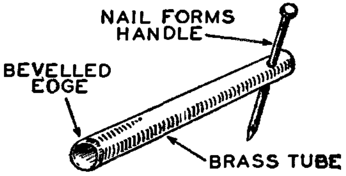
A strip of adhesive bandage will adhere firmly to clean glass and may be used as a label. Stick it on the glass and then write on the name with pen and ink.

You can write or print directly on glass with a pencil or crayon if you first roughen or frost the surface with a Carborundum stone. The label may be erased by scrubbing with a piece of soft rubber.

CORKS AND STOPPERS

Corks and rubber stoppers are a necessity to the chemist. They can be purchased in all sizes at drug-stores and chemical-supply houses. You can buy some sizes at the "dime" store. Several firms dealing in laboratory supplies sell an assortment of 150 corks containing 15 different sizes for 75 cents. Rubber stoppers in small quantities vary in price from 1 to 10 cents each, depending upon the size.

SET OF CORK BORERS



HOME-MADE CORK BORER

The only way to make a satisfactory hole in either a vegetable or rubber cork is to do it with a cork borer Cork borers are metal tubes with sharp cutting edges The sketch at the left is a set of three factory-made cork borers

The sketch at the right shows how to make cork borers by beveling the end of a brass tube with a file and providing a handle.

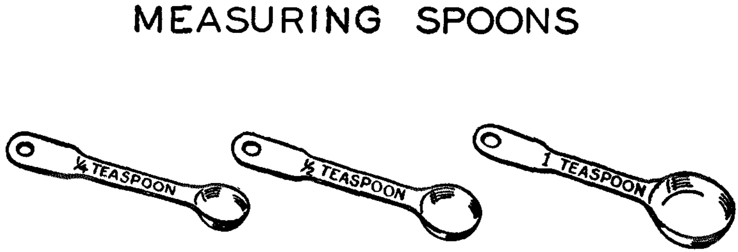
CORK BORERS

In order to perform some experiments, corks provided with one or two holes are necessary. When a set of cork borers is not handy, holes can be made with a red-hot nail. But the best method of boring a hole in a cork is to use a cork borer made just for that purpose. These can be used to make holes in rubber stoppers also.

A set of six borers costs about 65 cents. You can make your own out of some thin-walled brass or metal tubes by sharpening one end. File or grind one end to a bevel so as to form a sharp cutting edge. Drill the opposite end so that a nail can be slipped through to form a handle. When boring a cork, hold the cork against a piece of *soft* wood so that when the cutting edge of the tube comes through, it will not be dulled.

MEASURING CHEMICALS

Few amateur chemists, at least at the beginning of their careers, are able to possess a chemical balance or scale for weighing small amounts of chemicals. Consequently, in the experiments described in this book, the chemicals are measured by bulk or volume rather than by weight.

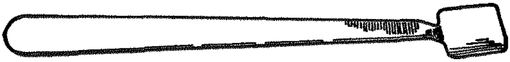

MEASURING SPOONS

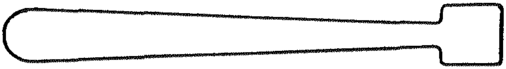
You can measure liquids and solids more easily and accurately with a measuring spoon than with an ordinary teaspoon. Measuring spoons, made either of pressed aluminum or molded plastic, are to be found on sale in the kitchenware department at the "dime" store. A set of four costs 10 cents.

A set of Bakelite or aluminum measuring spoons can be purchased for 10 cents at a "dime" store and will serve much better for chemicals than the ordinary metal teaspoon.

For small quantities of dry chemicals, make a "measure" according to the pattern shown in the accompanying illustration. Trace the outline with a piece of tissue-paper and paste it on a piece of cardboard or tin. Then cut it out. A tin measure will be more serviceable than one made of cardboard. A measure of this size holds about the same amount of a dry, powdered chemical as can be held on the end of the small blade of a penknife.

THE MEASURE





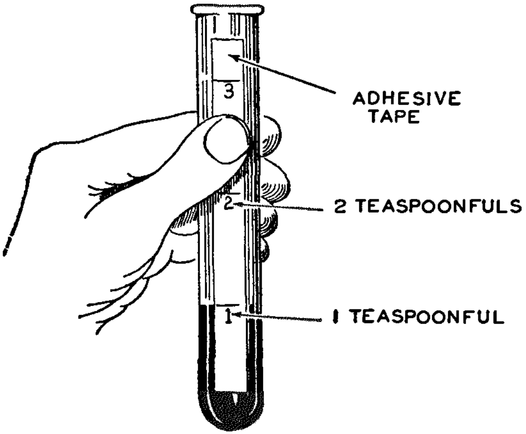
PATTERN

In many of the experiments described in the following pages, the "measure" shown in the accompanying illustration is used to measure or apportion small quantities of dry chemicals. Trace the pattern of the measure illustrated on tissue paper and paste it on a piece of Bristol board or tin. Then cut it out. A tin measure will be most serviceable.

A GRADUATE

In order to measure out small quantities of liquids when the experiment requires a certain number of teaspoonfuls, make a graduate. Stick a strip of adhesive tape lengthwise of a test-tube, starting at the bottom and running almost to the top. Measure out 1 teaspoonful of water and pour it into the tube. Make a pencil or ink mark on the adhesive tape to indicate the level to which 1 teaspoonful fills the tube. Add a second and a third teaspoonful, making a mark each time to indicate the level.

HOME-MADE GRADUATE



Acids and other liquids which attack metal should not be measured with a metal spoon. For this purpose use a graduate made from a test-tube. Stick a narrow strip of adhesive tape lengthwise to the tube. Put one measuring teaspoonful of water in the tube and make a mark on the adhesive tape to show the level reached. Add a second and third teaspoonful and make a mark each time to show the level. Pour out the water and use the tube as a graduate or measure.

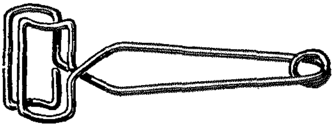
TEST-TUBES

The most used piece of glassware in a chemical laboratory is usually the test-tube. It is made in several sizes. The most useful size in the amateur's laboratory is from 5 to 6 inches long and from ⅝ to ⅞ of an inch in diameter. They cost from 25 to 30 cents per dozen. Some experiments spoil a tube for further use, but in most instances tubes can be used over and over again if properly washed. A test-tube brush (it costs 10 cents) is really useful, in fact, almost necessary, for cleaning tubes properly.

Every experiment should have a few test-tubes made of Pyrex glass. These withstand more shock and higher temperatures than the ordinary test-tube.

TEST-TUBE HOLDER

To avoid burning your fingers when heating a test-tube, you need a hold or clamp. You can bend a clamp out of wire or purchase one for 10 cents or less.



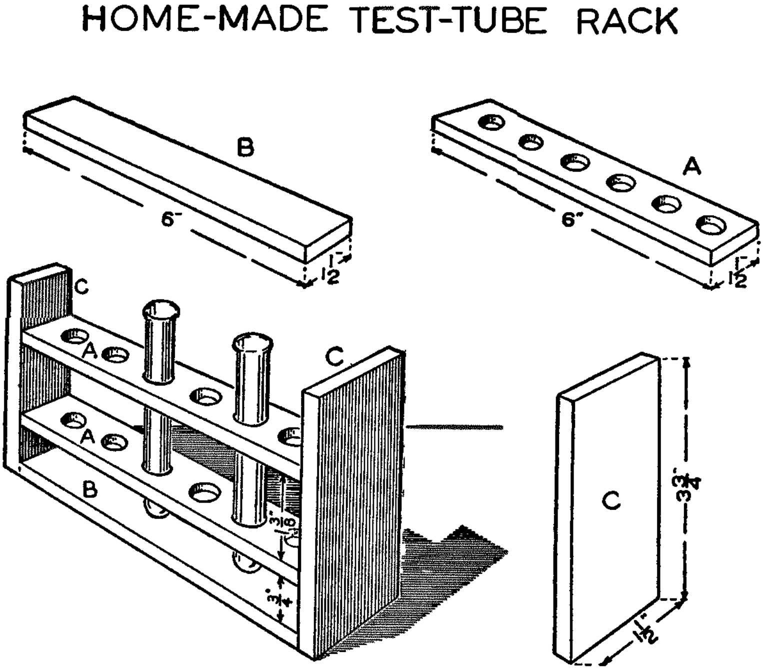
TEST-TUBE HOLDER

TEST-TUBE BRUSH

Here are two inexpensive pieces of equipment useful in any chemical laboratory. A test-tube holder is used when heating substances contained in a test-tube. A test-tube brush provides the only practical means of getting the inside of a test-tube really clean.

TEST-TUBE RACK

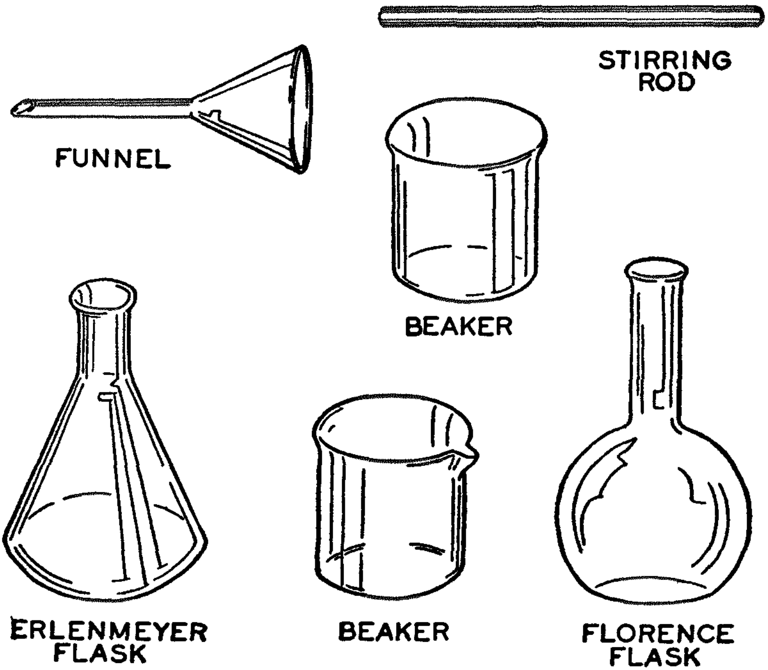
A test-tube rack, stand, or support (it is called by all three names) is useful for holding test-tubes. You can stand tubes in a glass tumbler or jar, but then it will not be as easy to observe their contents as it is when they are placed in a rack. A test-tube rack is quickly made by following the plan shown in the illustration.



Any one who can saw in a reasonably straight line and drill a few holes can make a test-tube stand or rack like that in the illustration above. Make two pieces like A and two like C. The diameter of the holes in A will depend upon the size of the test-tubes to be used in the stand.

FLASKS

It pays to buy flasks made of Pyrex glass. They are much stronger and withstand greater temperatures than ordinary glass flasks. Flasks have a great many uses and, if possible, several should be kept on hand. The Erlenmeyer type is probably the most satisfactory for general use. It has a larger base, consequently does not upset as easily and is more readily cleaned than a Florence flask. Flasks varying in size from 50 to 500 milliliters range in price from 20 to 50 cents.



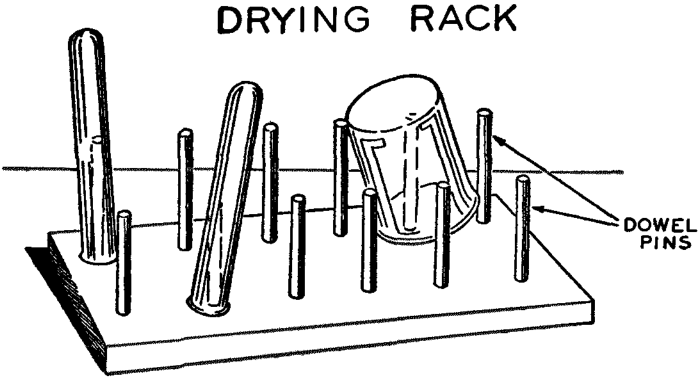
Here is some of the chemical glassware most frequently used which every young chemist should have in his laboratory.

BEAKERS

Three or four beakers of varying size are handy. Those with a lip or pour-out are most useful. If made of resistant or Pyrex glass, they will last longest.

DRYING RACK

Clean glassware is essential in a chemical laboratory. As soon as you are through using a test-tube, flask, or beaker, rinse it out, fill it with water, and let it stand until you are ready to wash it thoroughly. Then scrub it out, using a tube or bottle brush, rinse it, and set it upside down on a drying rack. A drying rack is made by setting several short pieces of dowel rod into holes in a wooden base as shown in the illustration.



When chemical glassware has been well washed it is rinsed in hot water and placed upside down on a rack to drain and dry. Wooden dowel rods driven into holes drilled in a piece of board make a handy drying rack.

STIRRING ROD

Shaking a test-tube containing a liquid and a solid usually causes the solid to dissolve more rapidly. But when chemicals are to be dissolved in a liquid contained in a beaker or flask, a stirring rod is handy. This is a glass rod, rounded at both ends. A glass "muddler," used for stirring iced beverages, makes a good stirring rod for a home laboratory. You can make one from a piece of glass tubing. Hold the ends of the tube in the flame of a Bunsen burner until they melt and close.

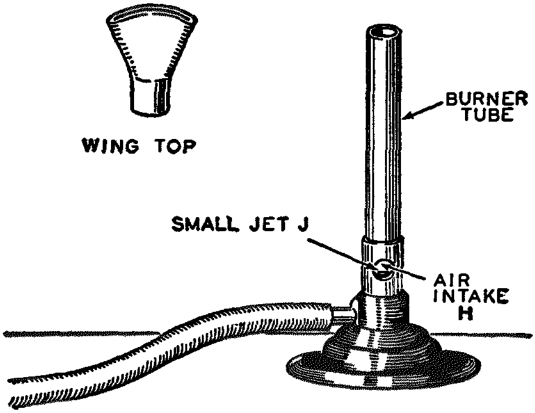
BUNSEN BURNER

Heat is required frequently in chemical experiments and processes. A flame that is not only very hot but that gives off no soot or carbon, is needed. Usually, chemical laboratories are equipped with Bunsen burners. A Bunsen burner is the ideal source of heat for most chemical experiments. It gives a clear, blue, sootless flame. It will burn illuminating gas, gasoline gas, or alcohol gas.

The burners on a gas stove are a form of Bunsen burner. The reason a hot, sootless flame is produced is that *air is mixed* with the gas *before* it is burned. A Bunsen burner mixes air with the gas before it is burned. This air furnishes enough oxygen to burn the carbon in the gas completely, and no soot is formed.

A Bunsen burner like that illustrated costs about 35 cents. It should be connected to a gas outlet preferably by a flexible metal gas tube 3 or 4 feet long. Flexible metal gas tubes made for connecting the small gas stoves called gas plates can

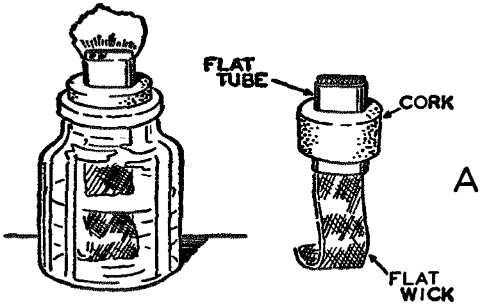
BUNSEN BURNER



A Bunsen burner is a more satisfactory source of heat than an alcohol lamp. It is hotter and can be better regulated. The principle of the Bunsen burner is employed in all gas heaters. When the burner is in use, gas enters the burner tube through the small jet (J) and mixes with a supply of air from the air intake (H). The air enters the intake holes at the base of the tube, because gas issues from the jet at high speed and draws along air with it. The wing top is fitted to the burner when a wide flame is needed for bending glass tubing.

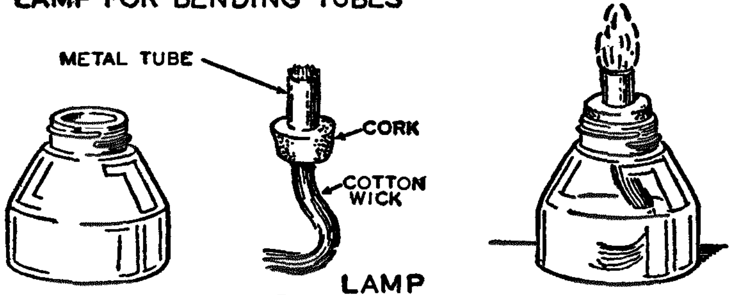
usually be obtained at a hardware store or at the gas company. Rubber tubing can be used in place of metal tubing wherever it is not "outlawed" by regulations of the local building or fire department. It should have thick walls and not be so small that it requires a great deal of stretching to slip over the gas cock and burner inlet. When rubber tubing gets old and brittle it should be thrown away.

Gas issues from the small jet, *J*, and rushes up the burner tube. The rush of gas up the tube draws in air through the adjustable air-hole, *H*.



ALCOHOL LAMPS

LAMP FOR BENDING TUBES



MUCILAGE BOTTLE LAMP FOR HEATING COMPLETE LAMP

An alcohol lamp is not nearly so desirable as a Bunsen burner in the laboratory, but sometimes is the only source of heat available.

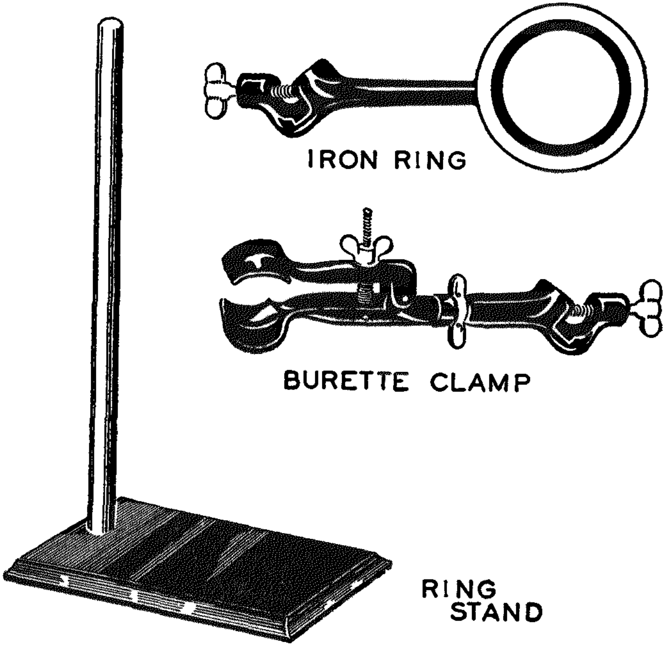
The illustration shows how to make an alcohol lamp out of a bottle, metal tube, wick and cork. A lamp with a round wick serves best for heating flasks and test-tubes. A lamp with a wide flat wick is used for heating glass tubing which is to be bent.

A wing top which will fit over the tip of the Bunsen burner and which will make a fan-shaped flame to bend glass tubing more easily costs 10 cents.

ALCOHOL LAMP

When gas is not available in your home laboratory, an alcohol lamp becomes necessary. One like that illustrated, made of heavy glass, complete with brass burner, wick, and ground-glass cap, costs about 40 cents.

You can make your own alcohol lamp. A low, squat bottle of the type used for mucilage or ink is ideal for that purpose because it will not tip over easily. The wick passes through a piece of metal tubing fitted in a hole in the cork. The wick can be one of the round cotton wicks made for the purpose and obtainable at a price of two for 5 cents, or it can be



A ring stand is one of the necessary pieces of apparatus in a chemical laboratory. The rings are made in several sizes and are used to support beakers, flasks and crucibles above a Bunsen burner or alcohol lamp. Burette clamps which attach to the stand are used to support burettes, condensers and test-tubes.

home-made and consist of several strands of white cotton cord.

When you try to bend glass tubing in a small round flame, the bend is usually too sharp and the tube becomes restricted.

For bending glass tubing with an alcohol lamp, a broad, flat wick is needed. A lamp to use the flat cotton wick sold for oil lanterns can be made by replacing the usual round tube through which the wick passes with a flat tube. Flatten a round metal tube to fit the wick, or bend a flat tube from a strip of sheet-tin. The hole in the cork must be a slot, made to fit the tube.

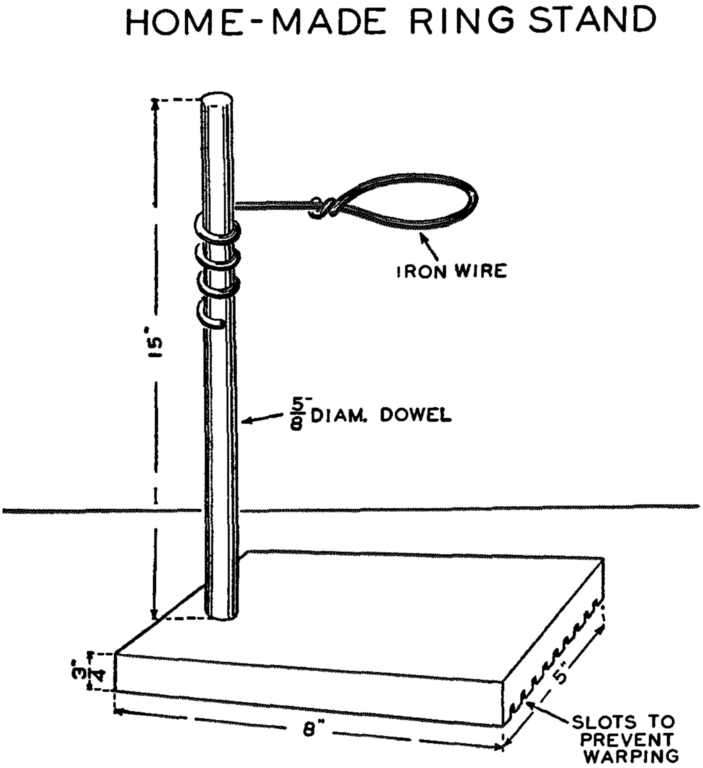
After filling an alcohol lamp, wait long enough before lighting it for any alcohol which may have spilled to evaporate.

Be certain that there is no alcohol on your fingers when you strike a match. It is an excellent idea to keep an alcohol lamp standing in a shallow pan of sand, both when it is lighted and also when not in use.

RING STAND

Some convenient means of holding test-tubes, flasks, and crucibles over a Bunsen burner or alcohol lamp is necessary in a laboratory. An iron ring stand, complete with two rings and a burette clamp, costs about $1.25. This is very useful equipment. The rings are used to hold flasks and crucibles and the burette clamp to hold test-tubes while heating.

You can make a ring stand from a piece of ½- or ⅝-inch wooden dowel 18 inches long, and a wooden base ¾ inch thick, 5 inches wide, and 8 inches long. The dowel rod fits snugly into a hole drilled in the base. If the under side is grooved by slotting with a saw or chisel, it will aid in preventing the base from warping. A coat of asphaltum paint will help make the wood resistant to chemicals.

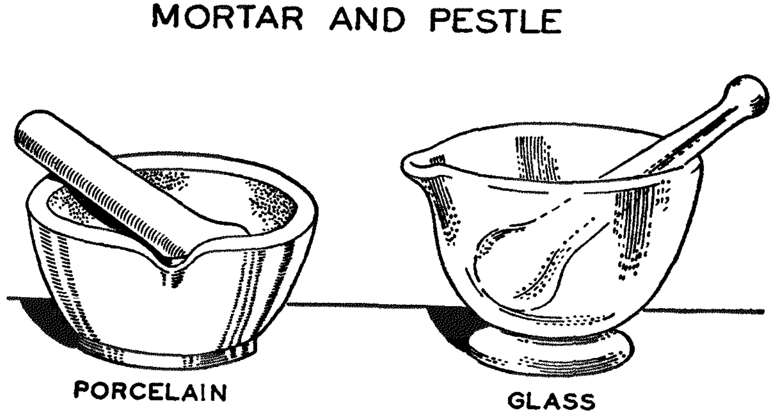


Commercial ring stands are made of iron and steel. A home-made substitute can be fashioned from wood and wire. The wire ring can be prevented from slipping down the dowel rod by a spring clothes-pin.

The rings for holding flasks, test-tubes, crucibles, etc., can be bent out of heavy wire. The wire used to make the wire coat-hangers which dry-cleaning establishments usually return with clothing is the right diameter and temper for this purpose.

MORTAR AND PESTLE

Often chemicals must be in pulverized form in order to mix and react with other substances. To pulverize small quantities of chemicals, you need a mortar and pestle. They are made in various sizes, of glass, stone, porcelain, iron, and brass. The 2-ounce size is large enough to prepare chemicals for any of the experiments described in this book. A small mortar and pestle made of heavy glass and with lip for pouring costs about 45 cents. A small mortar and pestle made of Coors porcelain costs about the same.



The mortar and pestle is useful equipment in any laboratory for grinding small quantities of chemicals to a fine powder. The small sizes either glass or porcelain cost about 45 cents.

USING A MORTAR AND PESTLE

Do not pulverize two different substances at the same time in a mortar and you will eliminate one cause of explosions. When two different substances are to be pulverized and mixed, they should be pulverized separately and mixed afterwards. Wash and dry the mortar and pestle each time that you use it for grinding a different substance.

GLASS FUNNEL

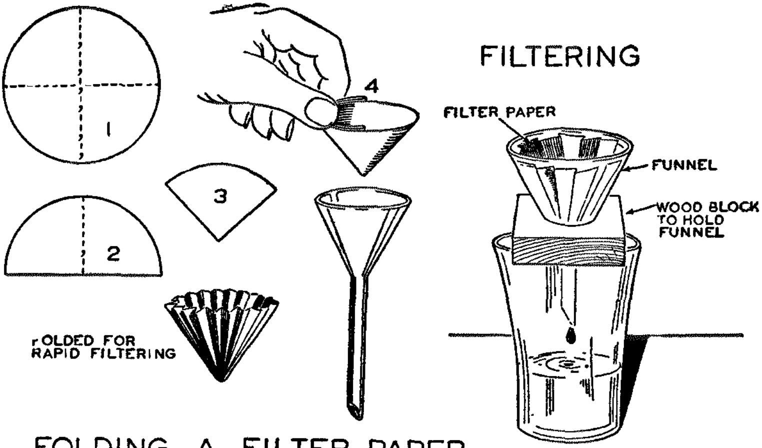
A funnel is not only useful for pouring liquids into narrow-necked bottles but is a necessity for filtering or separating liquids from solids. A glass funnel 2½ to 4 inches in diameter costs from 25 to 35 cents.

FILTER PAPER

Filter paper is a porous, white paper made for laboratory use and sold in packages containing 100 circular sheets. A package of 100 sheets, 3½ inches in diameter, usually costs 15 cents. Filter paper allows liquids to pass through but holds back the solids. Sheets are prepared for filtering so that they will fit in a funnel by folding them as shown in the illustration.

WIRE GAUZE

For certain experiments a "gentle" heat is necessary and the flame of a Bunsen burner or alcohol lamp must not be allowed to come into direct contact with the bottom of the beaker or flask. In that case, the flask or beaker should be supported on a piece of wire gauze. A piece of ordinary



FILTERING

FOLDING A FILTER PAPER

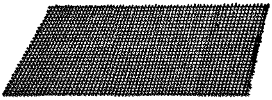
For ordinary purposes, fold a circular filter paper in half, then into quarters and open it, as shown in 1, 2, 3, 4. For more rapid filtering fold the paper into pleats. Bore a hole through a block of wood and use it to support the funnel when filtering.

window-screen about 4 inches square will serve. A sheet of asbestos paper laid on the screen will still further reduce the heat. You can buy a square piece of heavy wire gauze with an asbestos center at a laboratory-supply house for 8 to 10 cents. It will withstand the heat of a Bunsen burner longer than window-screen will.

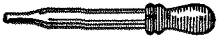
DEFLAGRATING SPOON

A deflagrating spoon is used to lower substances into a bottle of gas when they are to be burned therein. One can be purchased for 10 cents.

To make a deflagrating spoon, twist a wire around a metal bottle cap and bend it as shown in the illustration.



SCREEN



DROPPER

DEFLAGRATING SPOON

When heating flasks and beakers, it is frequently necessary to slip a small piece of wire screen underneath. This reduces the heat. Window screen may be used. To reduce the heat still further, put a piece of sheet asbestos between the screen and the glassware.

Twist a piece of stiff wire around a small metal bottle cap and you have a deflagrating spoon.

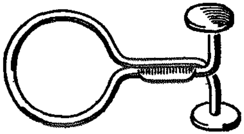
An ordinary medicine dropper is useful to the chemist.

PIPETTE

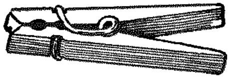
Pipette is the chemist's name for a medicine dropper. It is useful for measuring liquids by drops. A pipette is filled by inserting the open end in a liquid and squeezing the bulb. When the bulb is released, the liquid is drawn into the tube.

CLAMP OR PINCH-COCK

These are used to pinch or compress rubber tubing so as to shut off the passage of gases or liquids. The cost of a pinch-cock is usually 15 cents. You can use a spring clothes-pin for the same purpose.



MOHR'S PINCH-COCK



SPRING CLOTHES-PIN

Pinch-cocks are used to close rubber tubing by pressing the walls together. A spring clothes-pin can be used as a pinch-cock for small tubing.

CRUCIBLES

A crucible is needed for experiments which require the melting of metals and the heating of solids at temperatures which would soften glass. Clay crucibles can be used for melting metals in a stove or furnace. A clay crucible about 2½ inches in diameter and 4 inches deep costs about 25 cents.

Porcelain crucibles are provided with a cover so that the contents can be shielded from the air. They are more suitable than clay crucibles for heating small quantities of chemicals. Small porcelain crucibles varying in capacity from 15 to 50 milliliters (cubic centimeters) range in cost from 40 to 60 cents.

CHAPTER II

EXPERIMENTS WITH PRECIPITATES

SOMETIMES, when two clear solutions are mixed together, a solid material suddenly appears. This solid material may vary from a clear gelatinous substance or a white flaky material to an opaque, brilliantly colored mass. It is a new substance which did not exist in either of the solutions. It was created by the chemical reaction between the two liquids and is called a PRECIPITATE.

A precipitate is insoluble in the liquid in which it forms. Otherwise it would remain in solution and be invisible. Eventually it settles in the liquid.

A precipitate can be separated from the liquid in which it is suspended by filtering.

Mixing two solutions together so as to form a precipitate is one of the simplest of chemical experiments. Usually, the only equipment necessary is two test-tubes, the necessary chemicals, and some water. Making precipitates is an ideal way for the young experimenter to begin an acquaintance with chemistry. It is an interesting pastime. Some of the colors produced by the reactions are beautiful.

Throughout this book the chemicals required to perform each experiment are listed under the title of the experiment.

CRUCIBLES

CLAY

PORCELAIN

Crucibles will withstand more heat than glassware and are used in the chemical laboratory when it is necessary to heat substances to a high temperature for a prolonged period. When it is necessary to exclude the air during the heating process, a porcelain crucible with a snug fitting cover is utilized.

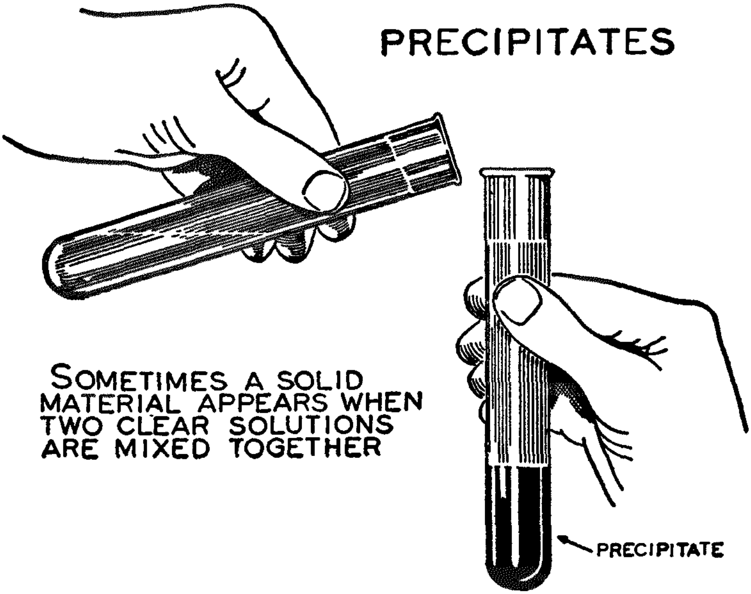
If you have a certain chemical on hand and wish to know what you can do with it, look it up in the index and you will be referred to those pages where experiments using that substance appear.

Azure Blue

COPPER SULFATE HOUSEHOLD AMMONIA

Put some small crystals of copper sulfate in a test-tube containing two or three teaspoonfuls of water. Shake the tube until the crystals dissolve and a solution of copper sulfate is formed. The color of this solution will be light green.

Add a few drops of ammonium hydroxide (ordinary household ammonia). An intense blue color will be formed. The color is produced by a complex copper-ammonia compound called an AMMINO SALT.



A precipitate is an insoluble solid that has been formed by chemical action in a liquid. The separation of an insoluble solid in a liquid is called precipitation. Precipitation is one of the chemical processes most frequently used in the laboratory. It is part of many manufacturing processes and is used in chemical analysis.

Prussian Blue

POTASSIUM FERRICYANIDE FERRIC CHLORIDE

Pour enough water into two clean test-tubes so that each is about one-quarter filled. Into one of the tubes drop one measure of potassium ferricyanide. Place one measure of ferric chloride crystals into the other tube. When these salts are completely dissolved, pour the solutions together into one tube. This forms FERRIC FERRICYANIDE and the brilliant blue color called Prussian blue.

Turnbull's Blue

POTASSIUM FERRICYANIDE FERROUS CHLORIDE

Pour enough water into two clean test-tubes so that each is about one-quarter filled. Into one of the tubes put one measure of potassium ferricyanide. Place one measure of ferrous chloride in the other. When these salts are completely dissolved, pour the potassium ferricyanide solution into the ferrous chloride solution.

FERROUS FERRICYANIDE is formed. It is deep-blue in color and is known as Turnbull's blue.

Orthosilicic Acid

SODIUM SILICATE SODIUM BISULFATE

Put 1 teaspoonful of sodium silicate solution in a test-tube. Add 1 teaspoonful of water and shake until thoroughly mixed. Dissolve ¼ teaspoonful of sodium bisulfate in a test-tube containing 2 teaspoonfuls of water. When these two solutions are mixed, the jelly-like substance known as ORTHOSILICIC ACID is formed.

There are several silicic acids. Many of the common minerals are silicates. Silicates are salts of the silicic acids. Clay and mica are acid orthosilicates.

Silicon Dioxide

ORTHOSILICIC ACID

Place some of the orthosilicic acid prepared in the previous experiment in a spoon and heat it over the flame of an alcohol lamp or a Bunsen burner. When the water has been driven out of the orthosilicic acid by the heat, the substance which remains in the spoon is SILICON DIOXIDE.

Silicon, and oxygen, are our most plentiful elements. Silicon is never found in the free condition. In combination it constitutes more than one-quarter of the crust of the earth.

The oxide of silicon is SAND. The substance obtained by heating orthosilicic acid is pure sand from the chemical standpoint.

Aluminum Silicate (White)

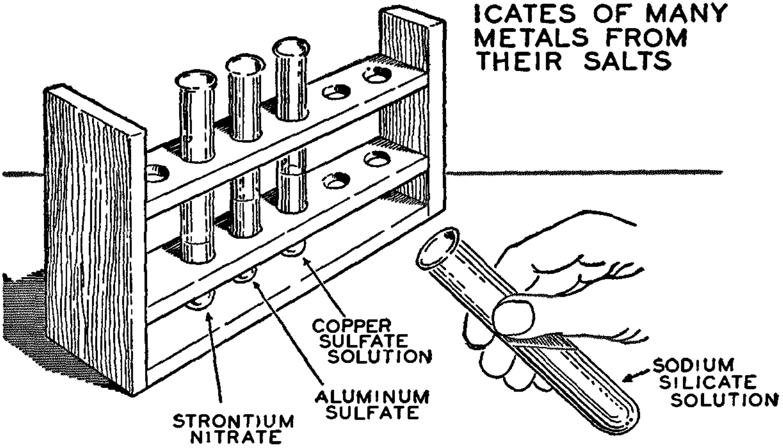
ALUMINUM SULFATE SODIUM SILICATE

Dissolve a small amount of aluminum sulfate in a test-tube half filled with water and then add several drops of sodium silicate solution. A white precipitate of ALUMINUM SILICATE will be formed.

Iron (Ferric) Silicate (Red)

FERRIC AMMONIUM SULFATE SODIUM SILICATE

Dissolve a small amount of ferric ammonium sulfate in a test-tube half filled with water and add a few drops of sodium silicate solution. A red precipitate of FERRIC SILICATE will be formed.



SODIUM SILICATE PRECIPITATES SILICATES OF MANY METALS FROM THEIR SALTS

The deepest wells have penetrated barely a mile and a half into the crust of the earth. We know little as to what is beyond this depth. That part of the solid portion of the earth with which we are familiar is three-quarters oxygen and silicon. One of the mineral forms of silicon is the silicates.

A mineral may be said to be any substance of definite constitution occurring in the ground. Rocks are mineral substances but not minerals. Many of the minerals are silicates. Among the well known mineral silicates are feldspar, mica, and kaolin.

Strontium Silicate (*White*)

STRONTIUM NITRATE SODIUM SILICATE

Dissolve a small amount of strontium nitrate in a test-tube half filled with water. Add a few drops of sodium silicate solution. A white precipitate of STRONTIUM SILICATE will be formed.

Copper Silicate (Blue)

COPPER SULFATE
SODIUM SILICATE

Dissolve some small crystals of copper sulfate in a test-tube half filled with water and then add a few drops of sodium silicate solution. A blue precipitate of COPPER SILICATE will be formed.

Cobalt Silicate (Blue)

COBALTOUS CHLORIDE
SODIUM SILICATE

Dissolve a few small crystals of cobaltous chloride in a test-tube half filled with water. Then add a few drops of sodium silicate solution and a blue precipitate of COBALT SILICATE will be formed.

Iron (Ferrous) Silicate (Green)

FERROUS AMMONIUM SULFATE
SODIUM SILICATE

Dissolve a small amount of ferrous ammonium sulfate in a test-tube half filled with water and then add three or four drops of sodium silicate solution. A green precipitate of FERROUS SILICATE will be formed.

Manganese Silicate (Pink)

MANGANESE SULFATE
SODIUM SILICATE

Dissolve a small amount of manganese sulfate in a test-tube half filled with water. Then add a few drops of sodium silicate solution. A pink precipitate of MANGANESE SILICATE will be formed.

Green Nickel Silicate

NICKEL AMMONIUM SULFATE SODIUM SILICATE

Dissolve a small amount of nickel ammonium sulfate in a test-tube half filled with water and then add a few drops of sodium silicate. A beautiful green precipitate of NICKEL SILICATE will be formed.

To Make Potassium Dichromate from Potassium Chromate

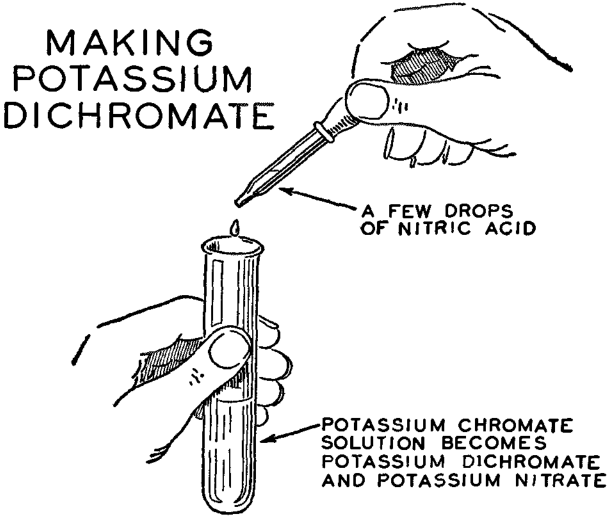
POTASSIUM CHROMATE NITRIC ACID

Dissolve some small crystals of potassium chromate in some water contained in a test-tube. Add slowly a few drops of nitric acid. The color will change from yellow to red. The red color indicates the presence of POTASSIUM DICHROMATE. Potassium nitrate is formed also. Since it forms a colorless solution, its presence is unseen.

Potassium chromate is used in dyeing, for tanning leather, and in the manufacture of pigments (chrome yellow, chrome red, etc.).

Potassium dichromate is used by the professional chemist in analytical testing. It is also the starting point for preparing many compounds of chromium.

A mixture of potassium dichromate and sulfuric acid is a powerful oxidizing agent and is often used for cleaning chemical glassware.



Potassium chromate forms a yellow solution. Potassium dichromate forms an orange red solution. Adding nitric acid slowly to a solution of potassium chromate will form potassium dichromate and potassium nitrate. The solution will change from yellow to red.

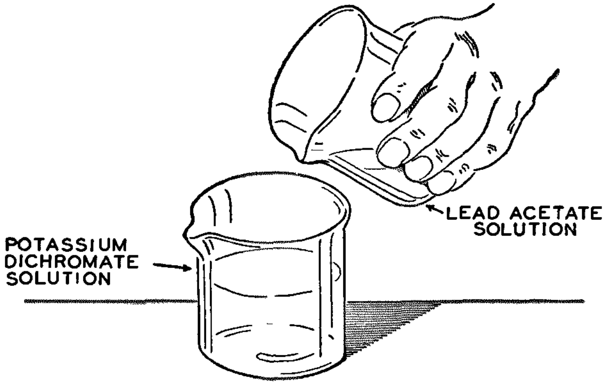
When potassium hydroxide is added to a solution of potassium dichromate, potassium chromate and water are formed. The solution changes from red to yellow.

To Convert Potassium Dichromate into Potassium Chromate

POTASSIUM DICHROMATE   POTASSIUM HYDROXIDE

Dissolve some small crystals of potassium dichromate in 2 or 3 teaspoonfuls of water contained in a test-tube. Dissolve some potassium hydroxide in water and add a little of this solution slowly to the potassium dichromate solution until the color of the latter changes from red to yellow. The reaction of the two substances forms POTASSIUM DICHROMATE and water.

TO MAKE CHROME YELLOW



Lead chromate is the chemist's name for the yellow pigment which painters call chrome yellow. Chromates are salts of chromic acid. They resemble sulfates in many ways. They crystallize alike

The chromates are all colored. In general chromium compounds are used for their color.

To Make Chrome Yellow

POTASSIUM DICHROMATE LEAD ACETATE

Chromium compounds give us some of our most useful colors for use in paints.

Dissolve 1 teaspoonful of potassium dichromate in water contained in a small beaker. In another beaker, dissolve 1 teaspoonful of lead acetate. Mix the solutions and a brilliant yellow precipitate will be produced. If undisturbed, this yellow powder will sink to the bottom of the beaker and the clear liquid can be poured off.

Pour clean water into the beaker. Stir it up so as to wash the powder. Let the powder sink once more. Pour off the water and spread the powder out on a filter paper so that it will dry. The yellow powder is LEAD CHROMATE, a useful pigment sold under the name of *chrome yellow.*

To Make Chrome Red

LEAD CHROMATE LIME OR CAUSTIC SODA

Place the lead chromate made in the last experiment in a beaker and cover it with water. Add a little lime or caustic soda and boil the mixture gently. The yellow lead chromate will change to orange and then to a brilliant red, becoming BASIC LEAD CHROMATE, another useful pigment.

To Make Chrome Green

POTASSIUM DICHROMATE AMMONIUM CHLORIDE

Chrome green is CHROMIC OXIDE. It is used for making green paint and for giving a green tint to glass. Chrome green can be made by heating a dry mixture of potassium dichromate and ammonium chloride. The common name of ammonium chloride is sal ammoniac.

Mix 1 teaspoonful of potassium dichromate with 1 teaspoonful of ammonium chloride. Place the mixture on a small piece of sheet-iron or the lid of a tin can and heat it over the flame of an alcohol lamp or Bunsen burner. Steam and nitrogen gas will be given off and there will remain a bulky green mass consisting principally of chrome green but containing some ammonium chloride and potassium chloride as impurities. These impurities can be removed by putting the green mass into hot water contained in a beaker and stirring thoroughly. The impurities will dissolve and the chrome green will settle to the bottom. Pour off the liquid, add more hot water, and stir. When the green pigment has settled, pour off the liquid once more and dry the residue.

To Make Blue and Green Verditer

CUPRIC NITRATE CALCIUM CARBONATE

Verditer is a light-blue pigment made by treating cupric nitrate with calcium carbonate.

Fill a test-tube or small beaker about one-third full of a solution of cupric nitrate made by dissolving cupric nitrate in water. Add to this solution some calcium carbonate (precipitated chalk) a little at a time. Keep stirring or shaking the mixture. A fine blue solid will separate out. When this has settled, pour in some clean water, stir it up, and let it settle again. Pour away the water again and you have left BLUE VERDITER. The washing process removed most of the calcium nitrate which was present.

To make GREEN VERDITER add some water to the blue verditer and boil it.

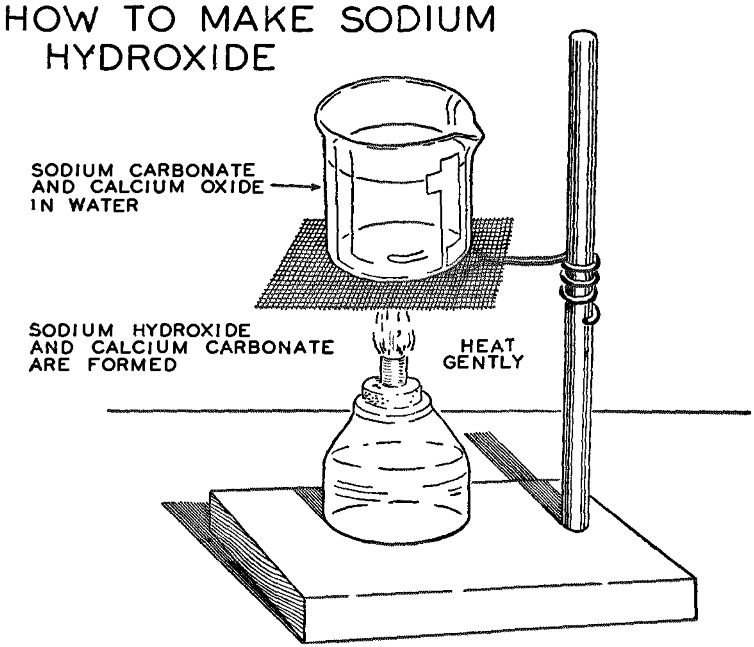
Making Strontium Sulfate

ALUMINUM SULFATE STRONTIUM NITRATE

Put 2 teaspoonfuls of water in a clean test-tube and add 4 measures of aluminum sulfate.

In a second test-tube also put 2 teaspoonfuls of water. Add to this 4 measures of strontium nitrate.

Pour one of these solutions into another and a white precipitate of STRONTIUM SULFATE will be formed. Aluminum nitrate is also formed by the reaction but it is soluble and does not appear unless the liquid is filtered so as to remove the strontium nitrate. Then, if the liquid is evaporated, crystals of aluminum nitrate are left as a residue.



Sodium hydroxide is used in the preparation of paper pulp, in the manufacture of soap and in many other chemical industries. It is produced commercially both by electrolysis of a solution of sodium chloride and by the action of slaked lime on sodium carbonate.

How to Make Sodium Hydroxide

SODIUM CARBONATE   CALCIUM OXIDE OR QUICKLIME

Sodium hydroxide, also known as caustic soda, is a strong alkali which is used in the manufacture of soap, in making paper pulp, and in many other chemical industries. A solution of sodium hydroxide has an exceedingly corrosive action upon the flesh. Even a weak solution feels slippery when rubbed between the fingers because it dissolves the skin, decomposing it into a slimy mass.

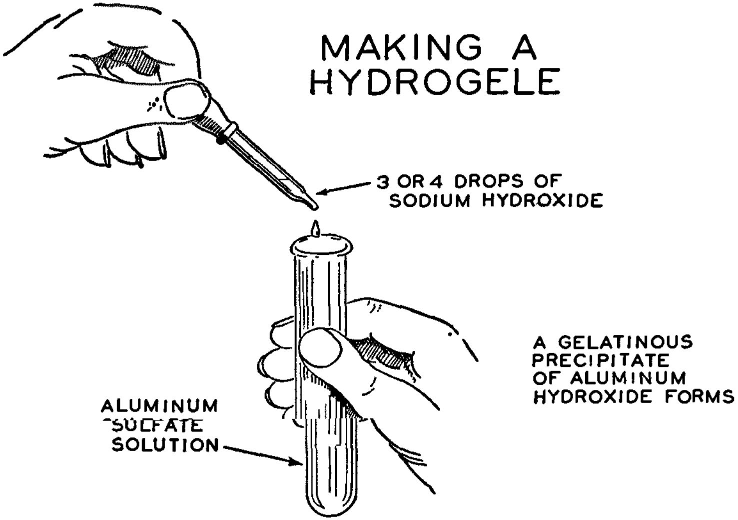
Dissolve 1 teaspoonful of sodium carbonate in a small beaker glass half filled with water. Then add a teaspoonful of calcium oxide or quicklime and stir well with a glass rod. Heat this mixture gently over a flame and then let it cool. The reaction produces sodium hydroxide and calcium carbonate. The sodium hydroxide is in solution but the calcium carbonate (chalk), being insoluble, forms a white precipitate. Let the precipitate settle and then pour off some of the clear liquid into a small bottle. Label it "Sodium Hydroxide Solution" and save it for further experiments. If you get any on your fingers, wash it off immediately.

To Make Aluminum Hydroxide

ALUMINUM SULFATE SODIUM HYDROXIDE

When an alkali is added to a solution of a salt of aluminum, a gelatinous precipitate of aluminum hydroxide is formed. This is a white gelatinous or jelly-like substance called in chemistry a HYDROGELE.

Dissolve 2 measures of aluminum sulfate in a test-tube half filled with water and add 3 or 4 drops of sodium hydroxide solution. The white precipitate which is formed is ALUMINUM HYDROXIDE.



The gelatinous precipitate of aluminum hydroxide is called a hydrogele. Aluminum hydroxide is a hydrogele which is often used in water purification. Aluminum sulfate and sodium hydroxide are added to the water. The aluminum hydroxide which is formed coagulates and settles out, entangling and carrying down refuse particles.

Aluminum hydroxide is a mild alkali and is used in medicine to remove excess hydrochloric acid from the stomach.

To Make Cobalt Hydroxide

COBALTOUS CHLORIDE SODIUM HYDROXIDE

Dissolve a few crystals of cobaltous chloride in a test-tube half filled with water. Add a few drops of sodium hydroxide solution. The blue precipitate which is formed is COBALTOUS HYDROXIDE.

To Make Ferrous Hydroxide

FERROUS AMMONIUM SULFATE SODIUM HYDROXIDE

Ferrous hydroxide forms as a precipitate when sodium or potassium hydroxide is added to a ferrous salt. Ferrous ammonium sulfate is a ferrous salt.

Dissolve a small amount of ferrous ammonium sulfate in water and add a few drops of sodium hydroxide solution. A green precipitate of FERROUS HYDROXIDE will form. This changes to a brown color when exposed to air and becomes *ferric* hydroxide, one of the compounds that composes rust.

To Make Ferric Hydroxide

FERRIC AMMONIUM SULFATE SODIUM HYDROXIDE

When a base (sodium hydroxide is a base) is added to a ferric salt (ferric ammonium sulfate is a ferric salt), a brown precipitate of ferric hydroxide is formed. In this form the ferric hydroxide is the jelly called a HYDROGELE by chemists. When it dries it becomes FERRIC OXIDE.

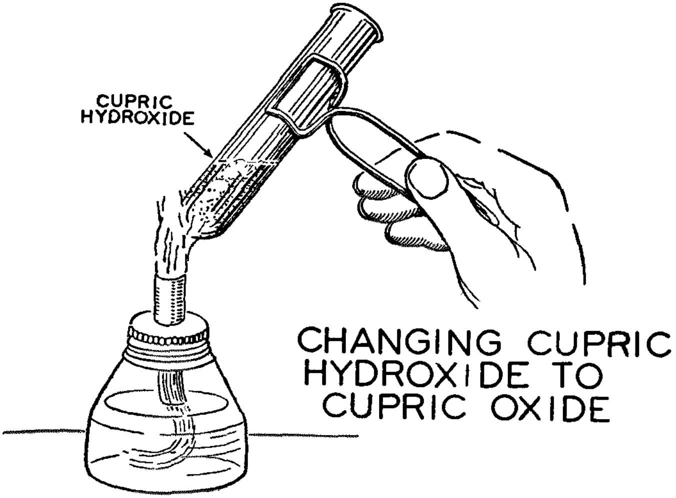
Dissolve one measure of ferric ammonium sulfate in a test-tube half full of water and then add a few drops of sodium hydroxide solution. The reddish-brown precipitate of FERRIC HYDROXIDE will be formed.

To Make Cupric Hydroxide

COPPER SULFATE SODIUM HYDROXIDE

When sodium or potassium hydroxide is added to a solution of a cupric salt, the gelatinous substance called cupric hydroxide is precipitated.

Dissolve 4 measures of copper sulfate in a small beaker glass half filled with water. While stirring this solution with a glass rod, slowly add sodium hydroxide solution. The gelatinous blue precipitate which forms is CUPRIC HYDROXIDE. In solution there will also be sodium sulfate and this can be removed by filtering. If too much sodium hydroxide is used, the precipitate will dissolve and the result will be a blue solution.



Cupric hydroxide is blue. Copper oxide is black. Cupric hydroxide is precipitated as a gelatinous substance when sodium or potassium hydroxide are added to a solution of a cupric salt. For example, sodium hydroxide added to a solution of copper sulfate will precipitate cupric hydroxide. When the mixture is boiled, the hydroxide loses water and forms a black hydrated cupric oxide.

Put some of the cupric hydroxide in a test-tube and heat it over a flame so as to drive off the water. The black substance which remains is CUPRIC OXIDE. If cupric oxide is strongly heated, it loses some of its oxygen and becomes CUPROUS OXIDE, which is red in color.

To Convert Cupric Hydroxide into Cupric Oxide

COPPER SULFATE
SODIUM HYDROXIDE

Precipitate some cupric hydroxide from a *cold* solution of copper sulfate by adding a small amount of sodium hydroxide as in the last experiment.

Heat this in the flame of a Bunsen burner or an alcohol lamp and notice the change from blue to black. The black substance is a mixture of cupric hydroxide and CUPRIC OXIDE.

To Make Nickel Hydroxide

NICKEL AMMONIUM SULFATE
SODIUM HYDROXIDE

Dissolve a few small crystals of nickel ammonium sulfate in a test-tube half filled with water. Add a few drops of sodium hydroxide solution. A blue-green precipitate of NICKEL HYDROXIDE will be formed.

An Experiment with Alum

ALUM
POTASSIUM HYDROXIDE

Aluminum sulfate will unite with potassium sulfate, sodium sulfate and ammonium sulfate to form complex chemical compounds called ALUMS. Potassium alum is the best known of these. Potassium alum is used in baking powder, in pigments and for purifying water.

Add 4 or 5 measures of powdered potassium alum to a test-tube half filled with water. When the alum has dissolved, add a few drops of potassium hydroxide solution. ALUMINUM HYDROXIDE will be precipitated. If you continue to add potassium hydroxide, the precipitate will dissolve.

Clarifying Water

LIME WATER CLAY ALUMINUM SULFATE

Aluminum hydroxide will clarify water.

Drop a small piece of clay into a bottle of water and shake until the water becomes cloudy. Put enough of this cloudy water into two test-tubes to half fill them. Set the tubes into the test-tube stand.

To one tube add about 20 drops of aluminum sulfate solution and then, after stirring with a glass rod so as to mix thoroughly, add 40 drops of *filtered* lime water. The gelatinous aluminum hydroxide formed by the reaction of these two compounds will settle toward the bottom of the tube carrying all the clay and foreign matter suspended in the water. Compare the contents of the tubes a short time later and notice that the water in the one without the aluminum hydroxide is still cloudy.

To Make Cobalt Phosphate

TRISODIUM PHOSPHATE COBALTOUS CHLORIDE

Phosphorus forms three chief acids called orthophosphoric acid, pyrophosphoric acid, and metaphosphoric acid. The salts of the phosphoric acids are called PHOSPHATES.

One of the common phosphates is trisodium phosphate. Solutions of trisodium phosphate are widely used in laundries, textile mills, and other industries where a cleansing agent is needed.

Most of the metallic phosphates are insoluble in water and consequently form a precipitate when a solution of one of their soluble salts is mixed with a soluble phosphate. A solution of trisodium phosphate will precipitate some of these metallic phosphates.

Dissolve 3 measures of trisodium phosphate in a test-tube containing 2 or 3 teaspoonfuls of water.

Dissolve 2 measures of cobaltous chloride in a test-tube containing 2 teaspoonfuls of water. Add some of the trisodium phosphate solution. A light-blue precipitate of COBALT PHOSPHATE will be formed.

To Make Ferric Phosphate

TRISODIUM PHOSPHATE FERRIC AMMONIUM SULFATE

Dissolve 4 measures of ferric ammonium sulfate in a test-tube containing 2 teaspoonfuls of water.

Add to this a few drops of a solution of trisodium phosphate made by dissolving 1 measureful in a teaspoonful of water.

The brownish-white precipitate which forms is FERRIC PHOSPHATE.

To Make Copper Phosphate

TRISODIUM PHOSPHATE COPPER SULFATE

Dissolve 3 measures of trisodium phosphate in a test-tube containing 2 or 3 teaspoonfuls of water.

In a second test-tube also containing 2 or 3 teaspoonfuls of water, dissolve 3 measures of copper sulfate.

Mix the two solutions together and a blue precipitate of COPPER PHOSPHATE will be formed.

To Make Aluminum Phosphate

TRISODIUM PHOSPHATE ALUMINUM SULFATE

Dissolve 3 measures of trisodium phosphate in a test-tube containing 2 or 3 teaspoonfuls of water.

In a second test-tube also containing 2 or 3 teaspoonfuls of water, dissolve 5 measures of aluminum sulfate.

Mix the two solutions together and a precipitate of ALUMINUM PHOSPHATE will be formed.

To Make Manganese Phosphate

TRISODIUM PHOSPHATE MANGANESE SULFATE

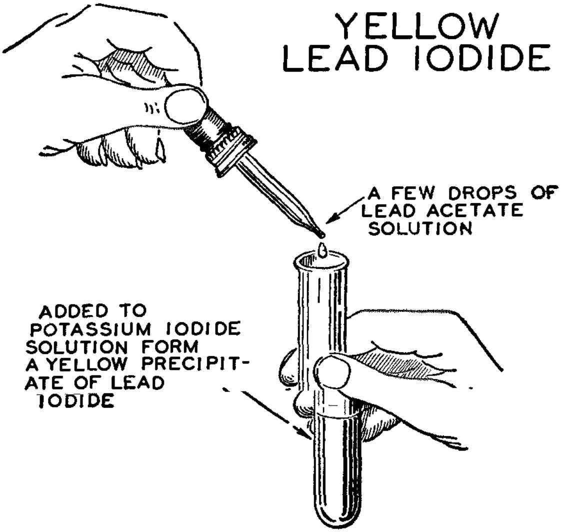
Dissolve 3 measures of trisodium phosphate in a test-tube containing 2 or 3 teaspoonfuls of water.

Put 6 measures of manganese sulfate in a test-tube containing 3 teaspoonfuls of water. It may be necessary to heat the water slightly in order to dissolve the manganese sulfate. When cool, mix the manganese sulfate solution with the trisodium phosphate solution and a faintly pink precipitate of MANGANESE PHOSPHATE will be formed.

Lead Iodide

POTASSIUM IODIDE LEAD ACETATE

Put 5 or 6 teaspoonfuls of water in a clean test-tube. Add a few drops of potassium iodide solution. Then add a few drops of a solution of lead acetate. A precipitate of brilliant yellow LEAD IODIDE will be formed.



Blue, white and green precipitates are quite common. This one is a brilliant yellow one.

Precipitating Copper Carbonate

COPPER SULFATE SODIUM CARBONATE

When carbon dioxide dissolves in water, it forms the weak acid known as CARBONIC ACID.

With each metal, carbonic acid forms two salts, an acid salt and a normal salt. Thus we have calcium bicarbonate and calcium carbonate; sodium acid carbonate and sodium carbonate. Carbonic acid easily breaks up. We are not likely to find bottles marked carbonic acid in laboratories.

The carbonates of many of the metals are insoluble in water. Consequently, they can be precipitated from solutions by passing carbon dioxide into the solution or by adding a soluble carbonate such as sodium carbonate.

Dissolve 2 measures of copper sulfate in a test-tube containing 2 teaspoonfuls of water. In another test-tube dissolve 2 measures of sodium carbonate in 1 teaspoonful of water.

Add a few drops of the sodium carbonate solution to the copper sulfate solution and a blue precipitate of COPPER CARBONATE will be formed.

Precipitating Manganous Carbonate

MANGANESE SULFATE SODIUM CARBONATE

Put 2 measures of manganese sulfate in a test-tube containing 2 teaspoonfuls of water. In another test-tube dissolve 2 measures of sodium carbonate in one teaspoonful of water.

Add a few drops of the sodium carbonate solution to the manganese sulfate solution and a white precipitate of MANGANOUS CARBONATE will be formed.

Pour the liquid and the precipitate into a filter so that the liquid runs through, leaving the manganous carbonate on the filter paper. As this dries it will become darker, changing to pink.

Precipitating Ferrous Carbonate

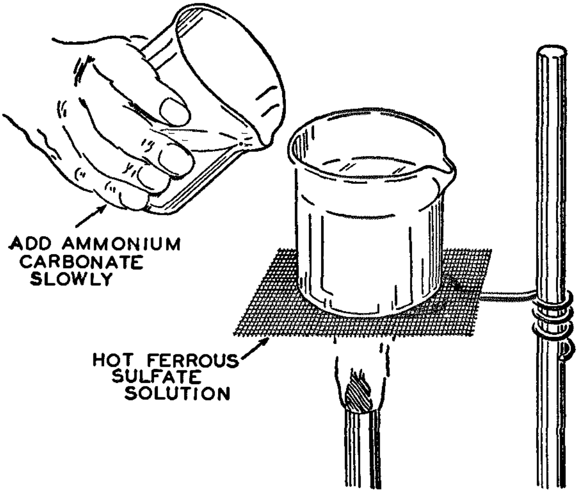
FERROUS SULFATE   AMMONIUM CARBONATE

Dissolve ¼ teaspoonful of ferrous sulfate in 8 teaspoonfuls of water contained in a small beaker.

Prepare a solution of ammonium carbonate by dissolving ¼ teaspoonful of ammonium carbonate in 6 to 8 teaspoonfuls of water.

Boil the ferrous sulfate solution gently over an alcohol

MAKING FERROUS CARBONATE

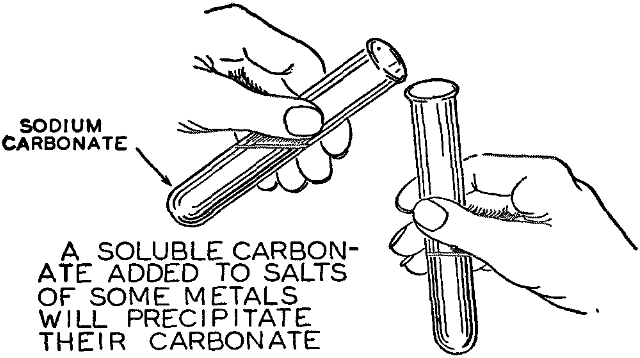


This experiment is intended principally for the amusement of the amateur chemist but it brings to light some interesting chemical facts. Ferrous carbonate is found in nature. It may also be made in a slightly hydrolized form by adding ammonium carbonate to a hot ferrous sulfate solution. The precipitate is white but soon darkens and becomes brown. This change in color to the subsequent formation of, first, ferrous hydroxide, and second, ferric hydroxide.

lamp or Bunsen burner and add some of the ammonium carbonate solution slowly.

A white precipitate of FERROUS CARBONATE will be formed. This precipitate will oxidize rapidly, becoming green, bluish-green, and finally RED FERRIC HYDROXIDE.

MAKING CARBONATES



The carbonates are salts of carbonic acid. With the exception of those of potassium, sodium, and ammonium, the carbonates are insoluble in water.

Precipitating Strontium Carbonate

STRONTIUM NITRATE SODIUM CARBONATE

Put 2 measures of strontium nitrate in a test-tube containing 2 teaspoonfuls of water. Shake until dissolved.

Make a solution of sodium carbonate by dissolving 2 measures of sodium carbonate in 1 teaspoonful of water.

Add some of the sodium carbonate solution to the strontium nitrate solution. The heavy white precipitate which is formed is STRONTIUM CARBONATE.

To Make White Lead

LEAD ACETATE SODIUM CARBONATE

White lead, or basic lead carbonate, is a heavy white powder which, when mixed with linseed-oil, is extensively used as a base for paints. It is produced by the reaction of lead with moisture and carbon dioxide.

Make a solution of lead acetate in a test-tube. Make a strong solution of sodium carbonate in a second test-tube.

Add the solution of sodium carbonate to the lead acetate as long as a precipitate is formed. The precipitate is WHITE LEAD.

Red and White Snow-Storm

SODIUM SILICATE SOLUTION STRONTIUM CHLORIDE
PHENOLPHTHALEIN SOLUTION

Put 5 teaspoonfuls of water in a tumbler or a small beaker and add 2 teaspoonfuls of sodium silicate solution. Mix thoroughly with a glass stirring rod.

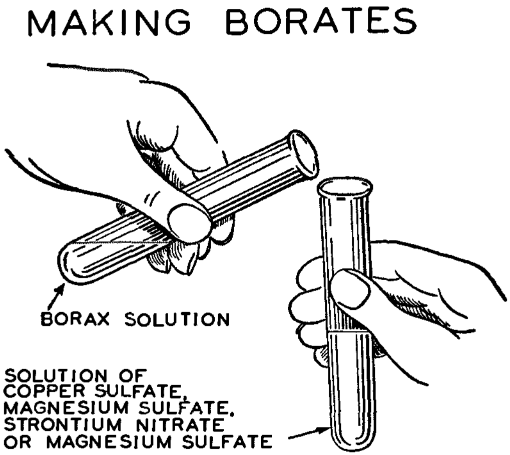
Next, put 2 teaspoonfuls of water in a test-tube. Add 6 measures of strontium chloride to this and shake until dissolved. Then pour a few drops of this strontium chloride solution into the sodium silicate solution. It will produce a white precipitate.

Add 3 or 4 drops of phenolphthalein solution to the remaining strontium chloride solution. Pour a few drops of this mixture into the sodium silicate solution and clots of red precipitate will appear among the white.

Making Borates

BORAX COPPER SULFATE ALUMINUM SULFATE
MAGNESIUM SULFATE MANGANESE SULFATE
STRONTIUM NITRATE

A solution of borax will react with solutions of many of the metallic salts. BORATES of the metals are thus produced. Solutions of copper, iron, aluminum, nickel, zinc chromium, calcium, strontium, magnesium, and manganese salts will react with a solution of borax to make borates of these metals.



Borates are salts of boric acid. Some borates are insoluble and may be precipitated by adding borax to certain solutions.

Dissolve some copper sulfate in a test-tube half filled with water and add a few drops of a strong borax solution, made by dissolving borax in water. A blue precipitate of copper borate will be formed.

You can substitute aluminum sulfate, magnesium sulfate, manganese sulfate, or strontium nitrate for copper sulfate and form borate of aluminum, magnesium, manganese, or strontium.

To Make Ferric Oxide

FERROUS SULFATE

When perfectly pure, the crystals of ferrous sulfate (also known as copperas and green vitriol) are transparent and of a bluish-green color. But in dry air they effloresce, that is, give up water, and become covered with a white incrustation. Through the absorption of oxygen, this subsequently changes to the rusty brown of ferric oxide.

Make a solution of ferrous sulfate in a small beaker and leave it exposed to the air for a few days. It will absorb oxygen from the air and, as the oxidation proceeds, the solution will gradually become yellow. After a while a brown precipitate of ferric oxide will appear.

CHAPTER III

EXPERIMENTS WITH SULFUR AND SOME OF ITS COMPOUNDS

SULFUR is an element of great importance in industry. A veritable mountain of this pale-yellow, brittle, crystalline solid, about 1½ million tons of it, is consumed each year in the United States alone. From sulfur is made sulfuric acid, gunpowder, matches, dyes, chemical sprays for trees and shrubs, colors, and photographic chemicals. It is used in disinfecting, bleaching, refrigeration, and in the process of vulcanizing rubber.

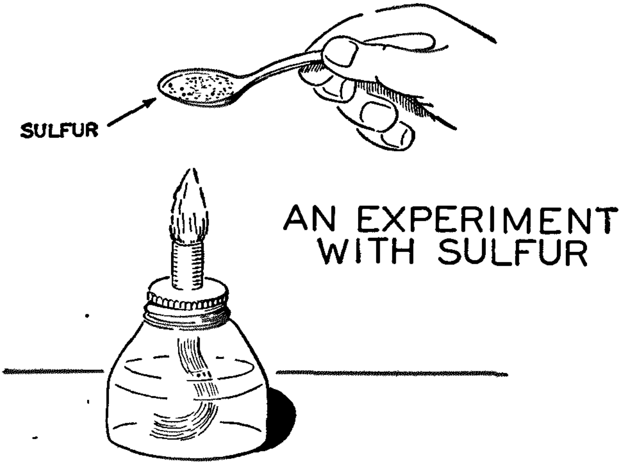
Sulfur itself is odorless and tasteless. The odor usually ascribed to sulfur, and its consequent bad reputation, is the odor of some of its compounds. Sulfur dioxide, formed when sulfur burns, has a strong suffocating odor. Hydrogen sulfide, a compound formed by a partnership between hydrogen and sulfur, has a very disagreeable odor like that of decaying organic matter.

Commercial sulfur is a crystalline substance which reaches the market in two forms, in lumps and as a fine, pale-yellow powder, insoluble in water. The powder (flowers of sulfur) is the most useful to the young experimenter.

Sulfur Melts Easily

POWDERED SULFUR

Put 1 teaspoonful of powdered sulfur in a dry test-tube and heat it slowly *over* the flame of an alcohol lamp or Bunsen burner. Apply the heat slowly and observe that the sulfur melts (238 degrees Fahrenheit) and forms a straw-colored liquid.



Heat gently a spoonful of powdered sulfur by holding it well above the flame of an alcohol lamp. It will melt and form a straw-colored liquid. If the heating is continued the color will change to brownish black. If held too near the flame the sulfur will become overheated and ignite. Strong suffocating fumes of sulfur dioxide will be formed. The burning sulfur may be extinguished by plunging it into water.

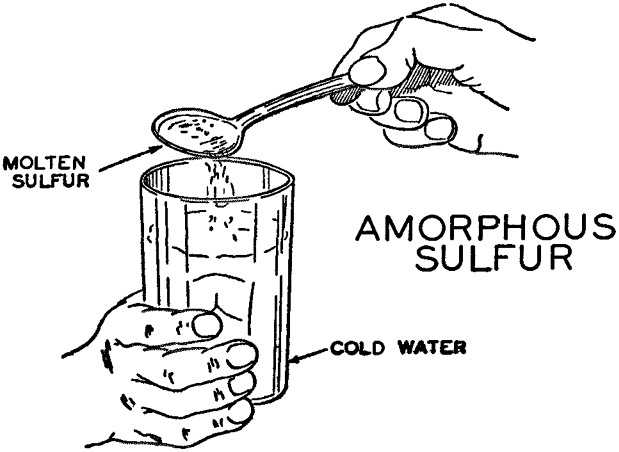
If the sulfur is melted in a test-tube instead of a spoon it will not catch fire as easily.

If the heating is continued gently, the pale-yellow molten sulfur will change in color and become brownish-black.

Amorphous Sulfur

POWDERED SULFUR

Melt a teaspoonful of sulfur in an old spoon well above a flame so that it will not catch fire. When the sulfur has melted, pour it quickly into a beaker filled with water. In a minute or so, take the sulfur out of the water and observe that it has become an elastic substance resembling rubber. It is non-crystalline and is called amorphous sulfur. Eventually it will change back into the ordinary crystalline form of sulfur.



Sulfur crystallizes in two distinct forms. Substances that crystallize in two distinct forms are called dimorphous. Sulfur can also be obtained in the amorphous or uncrystallized condition. Molten sulfur, suddenly cooled by pouring into cold water will be found to be soft and doughlike. After several days this elastic sulfur becomes hard and will be found to consist of rhombic sulfur crystals mixed with another variety of free sulfur without crystalline structure and called amorphous sulfur.

Sulfur Dioxide

An OUTDOOR Experiment

POWDERED SULFUR

Pour a teaspoonful of powdered sulfur into a small heap on a stone outdoors. Light the sulfur with a match and stand back. The sulfur will combine with oxygen in the atmosphere and form SULFUR DIOXIDE. You will smell its strong suffocating odor as the sulfur burns.

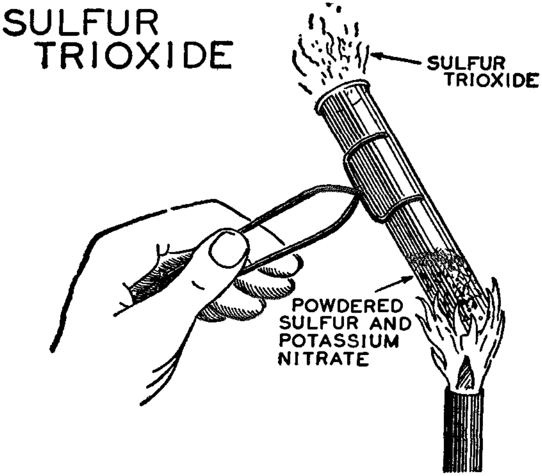
Sulfur dioxide is used to bleach wood pulp for paper-making, wool, and silk. It is also a disinfectant and a germicide (germ killer) as well as a bleach. Dried fruits such as peaches, raisins, cherries, etc., are sometimes bleached with sulfur dioxide.

How to Make Sulfur Trioxide

SULFUR POTASSIUM NITRATE

Sulfur trioxide is a white solid which dissolves in water to form sulfuric acid. Sulfuric acid is often called the "king of chemicals," for it is used directly or indirectly in every important industry. Sulfur trioxide is produced commercially in large quantities for making sulfuric acid by passing a mixture of sulfur dioxide and air over heated, finely divided platinum. The platinum acts as a catalyst. It brings about a reaction but does not play any other part in it and is not altered or consumed in any way.

Mix 1 measure of powdered sulfur with 1 measure of potassium nitrate on a piece of paper. Then put the mixture in a test-tube and heat very *gently.* White fumes will be given off. The fumes are SULFUR TRIOXIDE. If sulfur trioxide is dissolved in water, it forms SULFURIC ACID.



It is with oxygen that sulfur forms the compounds which bring it into greatest commercial use. The dioxide and trioxide are most important. When sulfur burns in the air it forms the well known suffocating sulfur dioxide. Under certain conditions, the dioxide combines with more oxygen forming sulfur trioxide. Heating a mixture of sulfur and potassium nitrate will produce sulfur trioxide.

Sulfur trioxide is a fluid at ordinary temperatures. It gives off strong, dense fumes when exposed to the air. With water, it reacts with great energy and forms sulfuric acid.

An Experiment with Sulfur

SODIUM HYDROXIDE POWDERED SULFUR

Put three teaspoonfuls of a solution of sodium hydroxide in a test-tube. Add ¼ teaspoonful of powdered sulfur. Using a test-tube holder, boil the solution gently for a few minutes in the flame of an alcohol lamp or Bunsen burner. The sulfur will disappear and the liquid will become dark-colored.

When the solution has cooled, add a few drops of hydrochloric acid to it. Sulfur will be precipitated and there will be a strong odor of hydrogen sulfide.

The Effect of Sulfuric Acid

SULFURIC ACID	SUGAR

Sulfuric acid attacks and destroys skin, leather, paper, cloth, and many other substances containing carbon, hydrogen, and oxygen.

The following experiment shows the effect of sulfuric acid on sugar and why this acid must always be handled carefully and with great respect.

Put 1 teaspoonful of sugar in a beaker glass. Spread it out evenly over the bottom of the beaker and add just enough water to cover it. Pour in a small amount of sulfuric acid (about 2 teaspoonfuls). The sugar will suddenly swell up into a great mass of porous carbon which may fill the whole beaker.

Hydrogen Sulfide

PARAFFIN	SULFUR

Shave enough paraffin from a white candle to fill a teaspoon when melted. Pour the molten paraffin into a test-tube and add 1 teaspoonful of powdered sulfur. Heat the mixture gently. In a minute or so remove it from the flame and smell the open end of the test-tube. The odor of rotten eggs is due to the formation of HYDROGEN SULFIDE.

The Combination of Iron and Sulfur

POWDERED SULFUR IRON FILINGS

Mix on a sheet of paper two teaspoonfuls of powdered sulfur with an equal amount of iron filings. Pour this mixture into a test-tube and heat it gently over an alcohol lamp or Bunsen burner.

You will soon notice a rather violent development of heat taking place in the contents of the tube. A glow will spread through the entire mass. When this has ceased, remove the test-tube from the flame and allow it to cool. Upon examination, you will find that the mass of sulfur and iron is a brittle, somewhat porous-looking black solid. Break it out of the test-tube and preserve it for experimenting with later.

This substance is known to chemists as FERROUS SULFIDE. It is sometimes called iron sulfide but ferrous sulfide is the more accurate term.

Hydrogen Sulfide

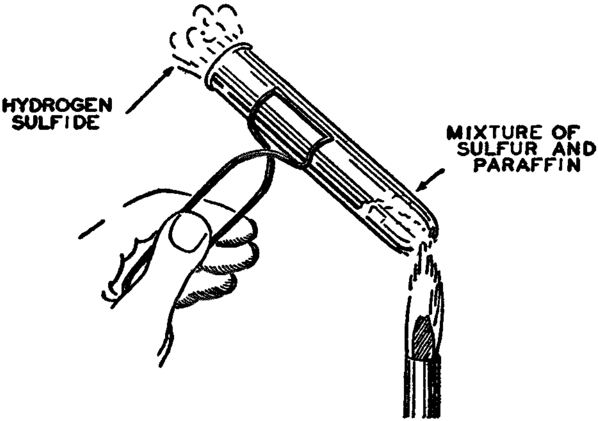
FERROUS SULFIDE SODIUM BISULFATE

Crush a piece of the ferrous sulfide prepared in last experiment to a powder and put a measureful in a test-tube. Add a measureful of sodium bisulfate and a few drops of water.

Warm this mixture over a flame for a moment. Then remove it from the heat and smell the mouth of the tube. It smells like rotten eggs. The odor is due to the gas called HYDROGEN SULFIDE, formed by the reaction between ferrous sulfide and sodium bisulfate.

Hydrogen sulfide is found in some mineral waters known as sulfur waters. It is produced in the decomposition of cabbages and eggs, the disagreeable odor of these substances being due largely to its presence. When concentrated or in large amounts the gas is poisonous but the small amount prepared in this experiment will not be harmful.

MAKING HYDROGEN SULFIDE



The preparation of hydrogen sulfide by treating ferrous sulfide with an acid is a classic experiment among high-school students and college freshmen. Heating a mixture of sulfur and paraffin will also produce the famous odoriferous gas.

The Preparation of Ammonium Sulfide

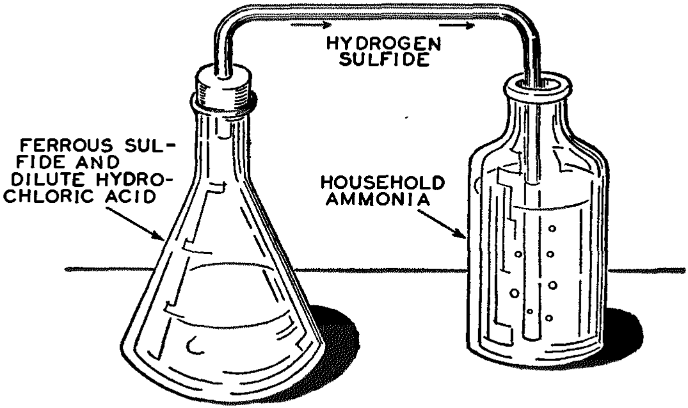
An OUTDOOR Experiment

FERROUS SULFIDE AMMONIUM HYDROXIDE SULFURIC ACID

Preparing ammonium sulfide makes a disagreeable odor. If you wish to avoid this you can purchase ammonium sulfide, already prepared, from a laboratory supply house.

Commercial laboratories and the laboratories in schools are usually provided with draft closets. These closets are connected to a flue or chimney which carries off fumes. Experiments which produce hydrogen sulfide, chlorine, and other obnoxious gases are carried out in the draft closet. Events can be observed through the glass door without inhaling any of the fumes.

PREPARING AMMONIUM SULFIDE



Ammonium sulfide is prepared by passing hydrogen sulfide into ammonium hydroxide. It is a colorless liquid having the disagreeable odor of ammonia and hydrogen sulfide. It soon changes color, becoming yellow. This change of color is due to a partial decomposition of the ammonium sulfide into ammonia, water, and sulfur.

A small amount of hydrogen sulfide in the atmosphere of a room is not harmful. It is merely a disagreeable odor. When a great deal of hydrogen sulfide is inhaled, it is poisonous. It then causes dizziness and headache—sometimes nausea.

It is perfectly safe to experiment with hydrogen sulfide if you make only a small quantity (use only ONE teaspoonful of ferrous sulfide) and experiment OUTDOORS or in a room with the windows wide open, so that any gas in the atmosphere of the room is greatly diluted.

Crush some ferrous sulfide into small pieces. Place a teaspoonful in an Erlenmeyer flask. The flask should be fitted with a one-hole stopper and a delivery tube which dips into a tall, narrow bottle. The delivery tube should be long enough so that it reaches almost to the bottom of the bottle. Pour about 6 ounces (¾ cupful) of strong household ammonia into the bottle. Pour a small amount of diluted hydrochloric or diluted sulfuric acid (storage-battery electrolyte will do) into the Erlenmeyer flask. Enough to cover the ferrous sulfide will be sufficient. Put the stopper and delivery tube in place. Hydrogen sulfide will be generated by the action of the acid on the ferrous sulfide and will pass through the delivery tube into the ammonia (ammonium hydroxide). As the gas bubbles up from the end of the tube it will be absorbed by the household ammonia and form AMMONIUM SULFIDE.

More acid should be added from time to time until all the ferrous sulfide has been consumed. Then pour the ammonium sulfide into a bottle having a tight-fitting cork. It is to be used for several experiments. The liquid smells strongly of ammonia and hydrogen sulfide. It can be used in place of hydrogen sulfide for performing several experiments and is not dangerous.

Tarnish

AMMONIUM SULFIDE

Brighten a cent and a dime with some metal polish. Wash and dry them. Put each coin in a test-tube and add a few drops of ammonium sulfide.

Do the bright coins lose their luster and become darkened?

The dulling and darkening of copper and silver articles in the home is called TARNISHING. Sulfur is usually responsible. Sulfur in the atmosphere forms a coating of copper sulfide or of silver sulfide on copper and silverware. The sulfur in the atmosphere of the average home comes from the chimney. Burning coal or oil releases sulfur fumes which reënter the house through doors and windows.

Copper (Cupric) Sulfide

COPPER SULFATE AMMONIUM SULFIDE

Make a solution of copper sulfate by dissolving a few crystals of that substance in water contained in a test-tube. Add a few drops of ammonium sulfide. A black precipitate will form. It is CUPRIC SULFIDE.

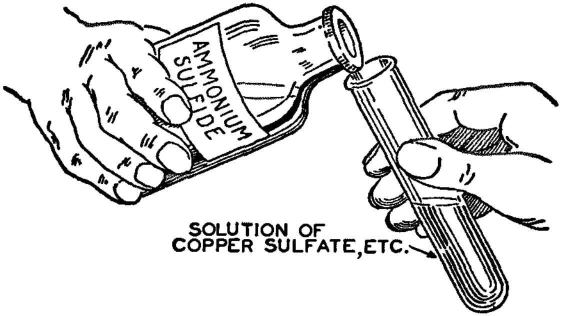
Several of the sulfides are insoluble in water and may be precipitated with hydrogen sulfide. For that reason, ammonium sulfide is much used in chemical analysis.

Lead Sulfide

LEAD ACETATE AMMONIUM SULFIDE

Dissolve some lead acetate in warm water contained in a test-tube and add a few drops of ammonium sulfide. A black precipitate of LEAD SULFIDE will be formed.

SEVERAL OF THE SULFIDES CAN BE PRECIPITATED



Some of the metallic sulfides are important. Many of them occur naturally as ores from which metals are extracted. Most of the sulfides are insoluble in water. The tarnishing of metals, particularly of silver, copper and brass is often due to the formation of sulfides on their surfaces. Some beautifully colored paint pigments are sulfides.

Zinc Sulfide

ZINC SULFATE AMMONIUM SULFIDE

Dissolve some zinc sulfate in water contained in a test-tube and add a few drops of ammonium sulfide. A *white* precipitate of ZINC SULFIDE will be formed. Zinc sulfide is the only familiar sulfide which is white.

Silver Sulfide

SILVER NITRATE AMMONIUM SULFIDE

Dissolve some small crystals of silver nitrate in water contained in a test-tube and add a few drops of ammonium sulfide. The black precipitate which forms is SILVER SULFIDE.

Copper (Cuprous) Sulfide

SULFUR SHEET-COPPER

Cuprous sulfide is produced when copper burns in sulfur vapor, or when copper is heated with sulfur. Cuprous sulfide differs from the cupric sulfide which forms as a precipitate when caustic soda or potash is added to a cold solution of copper sulfate. Two atoms of copper join with one atom of sulfur to form a molecule of cuprous sulfide, while in the cupric sulfide molecule there is one atom of copper and one atom of sulfur.

Heat a piece of sheet-copper in the flame of a Bunsen burner and, while it is red-hot, drop a small piece of sulfur on it. The reaction will form CUPROUS SULFIDE at the point of contact. Frequently a hole will be fused in the sheet.

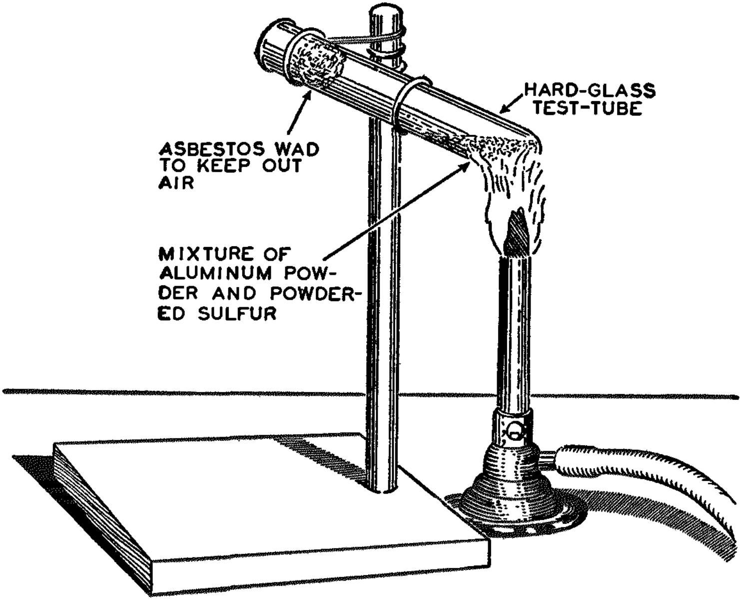
Aluminum Sulfide

ALUMINUM POWDER POWDERED SULFUR

Aluminum sulfide is formed when sulfur is heated with aluminum.

Mix 1 teaspoonful of aluminum powder with 1 teaspoonful of powdered sulfur. Place this mixture in a hard-glass test-tube. Put a wad of asbestos in the mouth of the test-tube to keep out the air and heat strongly in the flame of a Bunsen burner. ALUMINUM SULFIDE will be formed.

MAKING ALUMINUM SULFIDE



Aluminum sulfide is a grayish-black solid which is decomposed by water, producing aluminum hydroxide and hydrogen sulfide.

Aluminum Sulfide Reacts with Water

ALUMINUM SULFIDE WATER

Water decomposes aluminum sulfide and forms aluminum hydroxide and hydrogen sulfide. Atmospheric moisture decomposes this sulfide in a similar manner and consequently it must be kept in a tightly stoppered bottle.

When the aluminum sulfide prepared in the last experiment has cooled, pour some water in the test-tube. The gas which is evolved is HYDROGEN SULFIDE.

A Luminous Chemical

CALCIUM SULFIDE

The commercial (impure) sulfides of the alkaline earths (lime, magnesia, baryta, and strontia) appear luminous when placed in the dark after exposure to light. The pure sulfides do not have this property and the phosphorescence is therefore dependent upon the presence of some foreign substance.

Calcium sulfide is one of the ingredients of the luminous paint which has the peculiar property of glowing in the dark without giving off heat.

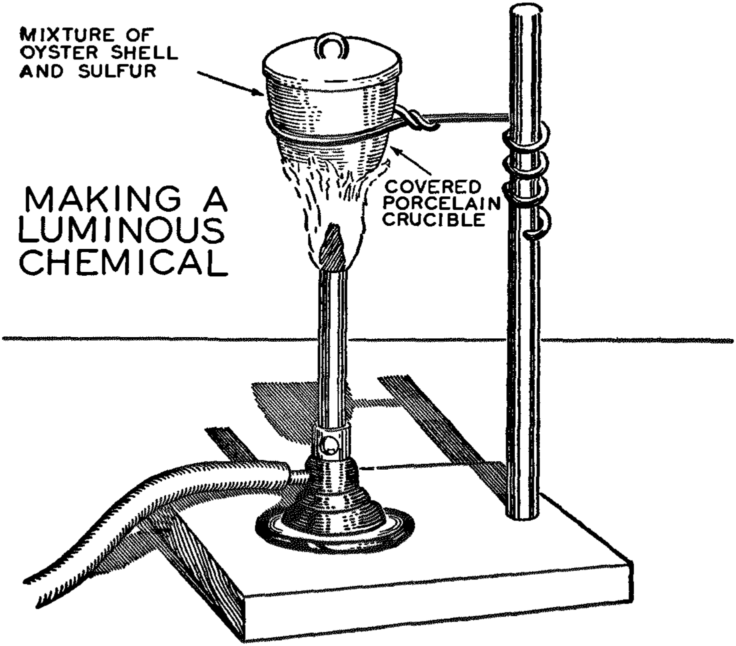
Fill a small bottle or test-tube with commercial luminous calcium sulfide. Cork it and expose to sunlight or a strong electric light for an hour or more. When taken into a dark room, it will glow with a faint, greenish light. The faint light will gradually diminish in intensity, but will regain its strength if again exposed to the sun or to a strong electric lamp.

Making Calcium Sulfide

OYSTER SHELLS SULFUR

You can make your own luminous calcium sulfide which will glow in the dark after exposure to the sun or after being held close to a strong electric light.

Grind some oyster shells to a fine powder in a mortar. Mix 2 teaspoonfuls of the powdered shell with 8 teaspoonfuls of sulfur. Place the mixture in a closed clay or porcelain crucible and heat red-hot for one-half hour or more. The result is CALCIUM SULFIDE. This should be ground to a powder



Some varieties of impure calcium sulfide, after they have been exposed to sunlight, shine in the dark. Barium sulfide also behaves in this way. Both these substances are used in making luminous paint.

A faintly luminous variety of calcium sulfide can be prepared by heating a mixture of oyster shell and sulfur.

in a mortar and placed in a tightly stoppered bottle. Moist air decomposes calcium sulfide.

To Make "Liver of Sulfur"

POWDERED SULFUR
POTASSIUM CARBONATE

Mix 1 teaspoonful of powdered sulfur with 1 teaspoonful of potassium carbonate and put the mixture in a clean test-tube. Heat gently over a flame until the effervescence ceases.

Then pour the molten mixture out of the tube onto a stone. Have an evaporating dish or porcelain crucible handy so that in case the mixture takes fire when it is poured out, you can extinguish it by covering and smothering it.

This fused mass, on account of its color, was once known to chemists as LIVER OF SULFUR. On exposure to the air for some time, it becomes greenish and moist and gives forth a small amount of hydrogen sulfide.

Hydrogen Sulfide from Liver of Sulfur

LIVER OF SULFUR SULFURIC ACID

Put some liver of sulfur in a test-tube, add water, and shake it up. A yellowish-green solution will be produced. If you add a few drops of diluted sulfuric acid to this, considerable hydrogen sulfide will be given out.

Antichlor

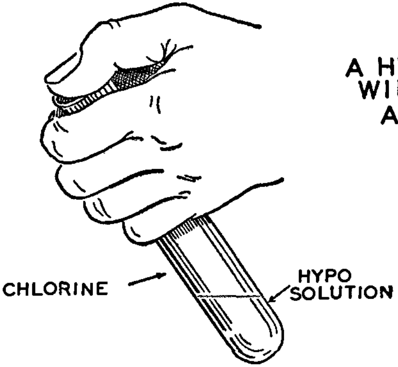
SODIUM THIOSULFATE CHLORINE

Make a solution of sodium thiosulfate (the "hypo" which photographers use) by adding 3 or 4 measures to a test-tube about half filled with water.

Pour some of this solution in a test-tube filled with chlorine (see page 105) and close the mouth of the tube with your thumb. Shake vigorously and observe that your thumb is sucked against the mouth of the tube.

A hypo solution is able to absorb and transform a very great amount of free chlorine almost immediately. When the chlorine in the tube was absorbed, it produced a partial vacuum. Sodium thiosulfate is employed in the textile business as ANTICHLOR, for the purpose of removing chlorine from bleached fabrics.

Notice that there is no longer any odor of chlorine in the tube.



A HYPO SOLUTION WILL QUICKLY ABSORB CHLORINE

Pour a little hypo solution (sodium thiosulfate) into a test-tube filled with chlorine gas. Close the mouth of the tube with your thumb. The hypo solution will absorb the chlorine so quickly that a partial vacuum will be formed. You can feel its suction against your thumb.

In the process of absorbing the chlorine, the hypo is decomposed into a good deal of free sulfur and sulfur dioxide. The milky appearance of the liquid in the tube is due to suspended sulfur. You can detect the sulfur dioxide by the odor in the tube. Other products of the reaction are sodium sulfate, sulfuric acid, and hydrochloric acid.

Sodium Iodide

TINCTURE OF IODINE

SODIUM THIOSULFATE

To a test-tube half filled with water add a few drops of tincture of iodine (the ordinary household iodine used for treating cuts). Observe the color imparted to the water.

Add a small amount of sodium thiosulfate (hypo) solution. The color will disappear immediately. The reaction between sodium thiosulfate and iodine produces SODIUM IODIDE and SODIUM TETRATHIONATE.

A solution of sodium thiosulfate may be used to remove iodine stains.

CHAPTER IV

EXPERIMENTS WITH OXYGEN AND OXYGEN COMPOUNDS

OXYGEN is the most abundant element. It makes up half our world. Oxygen is a gas having neither odor nor color. It is essential to both burning and breathing. Without it, neither activity can take place.

When another element combines with oxygen, it forms an OXIDE. The process of combining with oxygen is called OXIDATION. Oxidation is the correct term to use in referring to the chemical action occurring when any substance burns.

Oxygen forms about one-fifth of the volume of the atmosphere. It is therefore around us everywhere.

The easiest way for the experimenter to obtain oxygen is not from the atmosphere but from some of the compounds containing oxygen.

It can be obtained from water by passing an electric current through it under suitable conditions.

This method and the apparatus necessary are described in *Things a Boy Can Do with Chemistry** and *Things a Boy Can Do with Electrochemistry.**

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* By Alfred Morgan (New York, D. Appleton-Century Company, 1940).

Making Oxygen from an Oxide

MERCURIC OXIDE

Place ½ level teaspoonful of mercuric oxide on a small sheet of paper, folded down the middle. Slip the powder from the sheet of paper into a clean, dry test-tube. Hold the test-tube in an almost horizontal position and heat its lower end. Keep it moving back and forth in the flame. The tube may crack unless you are careful.

Notice that the red powder changes its color and becomes black.

After heating for a minute or two, thrust into the tube a splinter of wood which has a spark on the end. Repeat this test until the glowing splinter bursts into flame and burns brilliantly. This happens because the air has been pushed out of the tube by OXYGEN.

The oxygen was driven forth from the mercuric oxide by heat. One of the most characteristic properties of oxygen is its vigorous support of combustion.

Remove the test-tube from the flame and carefully examine the upper portion of the inner walls where the tube has not been very hot. It will appear like a mirror. It is coated with a thin layer of mercury.

Mercuric oxide consists of mercury and oxygen. When heated, it yields MERCURY and OXYGEN. The mercury is vaporized by the heat but condenses on the cooler parts of the tube and appears there as a mirror.

Invert the test-tube over a clean sheet of paper so that any undecomposed mercuric oxide remaining may fall out. It may be necessary to scrape it out with a small stick. Notice that the black powder changes its color as it cools and becomes red again. You may find some tiny globules of mercury mixed with the red powder.

Mercuric oxide is made by heating mercury and air or mercury and oxygen under proper conditions. Oxygen was discovered by a chemist named Priestly. In Priestly's day mercuric oxide was called *mercurius calcinatus*. These are Latin words and mean "mercury that has been heated in contact with air."

Burning Steel in Oxygen

MERCURIC OXIDE	STEEL WOOL

Place a teaspoonful of mercuric oxide in a large test-tube. Hold the tube in an almost horizontal position and heat the lower end in the flame of an alcohol lamp or Bunsen burner. Keep the tube moving back and forth so that it does not crack.

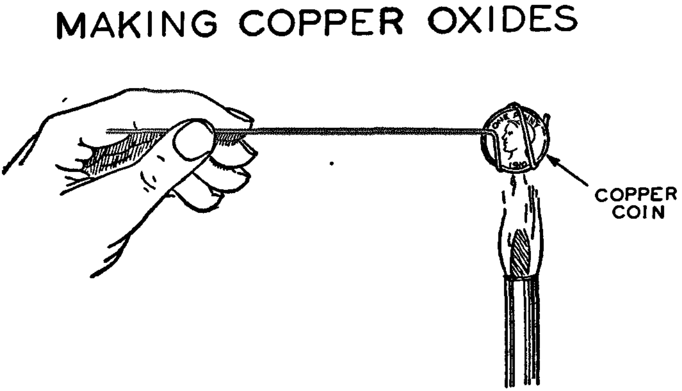
Bend a small hook in the end of an iron wire. Fasten a small wad of fine steel wool in the hooked end of the wire. Heat the steel wool until it is red-hot and plunge it *immediately* into the test-tube of oxygen. The steel wool will burn brilliantly.

When you withdraw the wire from the tube, notice a fine reddish-brown powder deposited on the sides of the tube and on the wire and any steel wool remaining. Steel is principally iron. When iron burns it forms IRON OXIDE. The reddish-brown powder is iron oxide.

Making Copper Oxides

COPPER COIN OR SHEET COPPER

There are two oxides of copper. One is red and the other is black. The red oxide is called CUPROUS oxide and the black oxide is called CUPRIC oxide. There is twice as much oxygen combined with the copper in cupric oxide as there is in cuprous oxide.



When a copper coin is heated in the flame of a Bunsen burner until it becomes black, two different oxides of copper will have been formed. One is called cuprous oxide and the other cupric oxide. Cuprous oxide is red and is produced by the oxidation of copper at a gentle heat Cupric oxide is a black substance formed by heating copper to redness in contact with the air. Heat a copper coin to redness until black cupric oxide forms and it will be found that underneath the black cupric oxide is an underlying layer of red cuprous oxide.

Brighten a copper coin or a small piece of sheet-copper by rubbing it with metal polish. Twist a wire around the coin so that the wire serves as a handle and the coin may thus be held in the flame of a Bunsen burner. If the coin is moved to and fro in the flame, a beautiful play of iridescent colors will appear upon the surface of the copper.

If, at this stage, the hot coin is plunged suddenly into water, it will become a dull-red color. The color is due to a thin coating of red CUPROUS OXIDE.

Heat the coin again, this time holding it in the upper or OXIDIZING portion of the flame. It will soon become covered with a layer of black CUPRIC OXIDE.

When a thick coating of the black oxide has been formed, plunge the hot coin into water again. The black coating of cupric oxide will scale off in spots, exposing a thin film of the dull-red cuprous oxide firmly adhering to the metal underneath.

Reducing Cupric Oxide

CUPRIC OXIDE

CHARCOAL

The opposite of oxidation, or forming an oxide, is the chemical action called REDUCTION. Heating mercuric oxide so that the oxygen and mercury in that compound are separated, as in the first experiment in this chapter, is reduction.

Some oxides can not be reduced as easily as mercuric oxide. For example, black copper oxide is not reduced merely by heating as is mercuric oxide. It is necessary to use a REDUCING AGENT. Charcoal (carbon) is a useful reducing agent. When mixed with black oxide of copper and heated, charcoal will take the oxygen of copper oxide away from the copper.

Mix 1 teaspoonful of copper oxide with 1 teaspoonful of powdered charcoal and place the mixture in a hard-glass test-tube. Grasp the tube in a test-tube holder or place it in a burette clamp on the ring stand and heat the mixture strongly in the flame of a Bunsen burner.

The carbon of the charcoal will unite with the oxygen of the copper oxide and escape in the form of the gas called CARBON DIOXIDE. The gas in soda-water and other sparkling beverages is carbon dioxide.

Metallic copper will remain in the tube. The carbon in the charcoal has *reduced* the copper oxide and the copper oxide has *oxidized* the carbon.

The separation of metals from their ores (metallurgy) is often a chemical process which utilizes the reducing power of carbon.

Making Oxide of Tin from Tin-foil

CUPRIC NITRATE TIN-FOIL

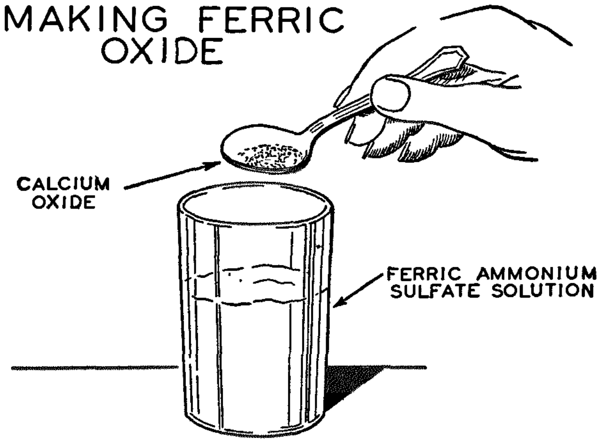
Heating or burning substances with flame is not the only method of forming the oxide.

The oxygen in certain compounds may be used to oxidize some substances. Cupric nitrate, in addition to copper and nitrogen, also contains oxygen.

A few years ago the foil used to wrap candies, chewing-gum, tobacco, etc., was tin-foil. Now a great deal of it is *aluminum* foil. Tin-foil has a slightly different color from aluminum foil and also a different crackle or "cry," as the sound made when the foil is bent or crinkles is called.

If you can distinguish between tin-foil and aluminum foil and select a piece of tin-foil, you can perform an interesting experiment.

Place some cupric nitrate crystals on a piece of tin-foil. Moisten them with a drop of water and quickly wrap them in the foil. Press the foil tightly against the crystals and drop the "package" upon a stone. Soon it will bubble and smoke. The tin is oxidized by the nitrate and converted into an OXIDE OF TIN.



Ferric oxide is found in nature in the form of a mineral called hematite. Rouge, Indian red, Venetian red, and metallic brown are ferric oxide in varying degrees of coarseness and purity. Large quantities of ferric oxide are manufactured from the ferrous sulfate obtained in cleaning ironware with sulfuric acid in preparation for tinning and galvanizing.

Making Ferric Oxide

FERRIC AMMONIUM SULFATE CALCIUM OXIDE

Compounds of iron are called *ferrous* and *ferric,* depending upon the relative amount of IRON which they contain. Ferrous compounds contain MORE iron in proportion to the other elements than ferric compounds do.

Put ½ teaspoonful of ferric ammonium sulfate in a tumbler half filled with warm water and stir with a glass rod until dissolved. Then add ½ teaspoonful of calcium oxide or quicklime and stir thoroughly. A solid brown precipitate will form. This is FERRIC OXIDE.

Here is what happened. The calcium of the calcium oxide replaced the iron in the ferric ammonium sulfate, forming calcium sulfate and ferric oxide. The ferric oxide is insoluble and can be separated by filtering. The calcium sulfate, being soluble, remains in solution and passes through the filter.

To Make Red Oxide of Lead

LEAD MONOXIDE (LITHARGE) POTASSIUM CHLORATE

Before performing this experiment read the precautions regarding potassium chlorate given in the list of chemicals in Chapter XII.

Mix 1 teaspoonful of lead monoxide with ½ teaspoonful of potassium chlorate. Put this mixture in a test-tube and heat it in a Bunsen-burner flame. The yellowish mixture will change into a red powder. This is MINIUM or RED OXIDE OF LEAD

Red lead is used in glass-making and, mixed with oil, as a paint for ironwork.

To Make Lead Dioxide

RED OXIDE OF LEAD NITRIC ACID

Put ½ teaspoonful of red oxide of lead in a test-tube. Add some water and then a little nitric acid. Heat the mixture gently for a few minutes. The red lead will be changed into a brown oxide of lead called LEAD DIOXIDE.

Oxidizing Sulfur

LEAD PEROXIDE FLOWERS OF SULFUR

Lead peroxide is a strong oxidizing agent and sulfur is always looking for oxygen.

Mix ½ teaspoonful of lead oxide with an equal amount of powdered sulfur, or flowers of sulfur. Rub some of this mixture on a warm surface with a stick. It will take fire and burn, giving out the obnoxious fumes of burning sulfur. When sulfur burns, it forms SULFUR DIOXIDE.

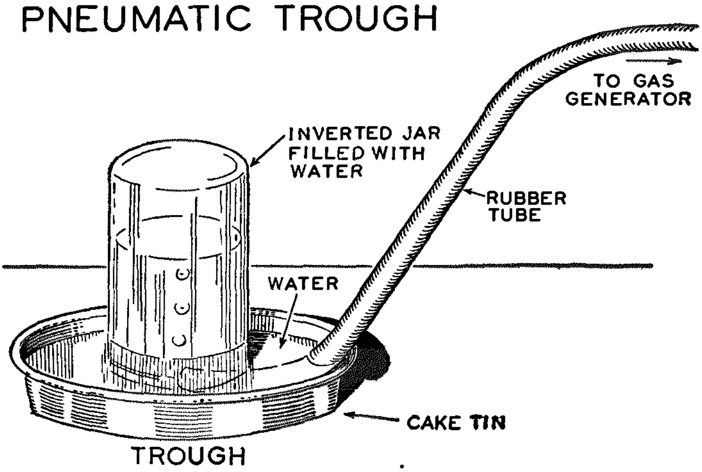
Oxygen from a Mouth-Wash

SODIUM PERBORATE MANGANESE SULFATE

Sodium perborate is to be found in many bathroom cabinets. It is used in the laundry for bleaching linen and cotton and for hygienic purposes. Some of the powders sold under trade-marked names as mouth-washes are merely flavored sodium perborate.

Sodium perborate is an oxidizing agent. It can be made to give up some of its oxygen by manganese sulfate. The manganese sulfate does not enter into the reaction. It merely acts as a CATALYST and helps it along.

Make a solution of sodium perborate in a test-tube and add several drops of a manganese sulfate solution. OXYGEN gas will escape.



The pneumatic trough is used for collecting certain gases. It can not be used for gases which readily dissolve in water.

A pneumatic trough is easily made. An inverted bottle or jar filled with water is placed mouth downward in a shallow pan or cake tin partly filled with water. The water in the pan seals the mouth of the bottle or jar and prevents air from entering and water from escaping. The end of a small rubber tube connected to the gas generator is placed under the bottle or jar. Gas escaping from the tube rises and displaces the water.

In order to place a jar in the trough fill it with water and lay a sheet of glass over its mouth. Invert the jar, lower it into the trough and when the mouth is under water remove the glass. To remove the jar when it has been filled with gas, slide a sheet of glass over its mouth and turn it rightside up.

HOW TO MAKE A PNEUMATIC TROUGH

A pneumatic trough provides a simple means of filling jars and bottles with gases which are not quickly absorbed by water. Carbon dioxide, hydrogen, oxygen, and nitrogen may be collected in this manner.

The principle of the pneumatic trough is simple. If a jar or a bottle is filled with water and inverted in a basin of water so that the mouth of the jar or bottle is below the surface of the water in the basin, the water will not run out of the jar or bottle. The pressure of the atmosphere on the surface of the water in the basin holds it in.

If bubbles of gas are allowed to rise in a bottle or jar filled with water and inverted in a basin, the water will be displaced and the bottle or jar become filled with gas.

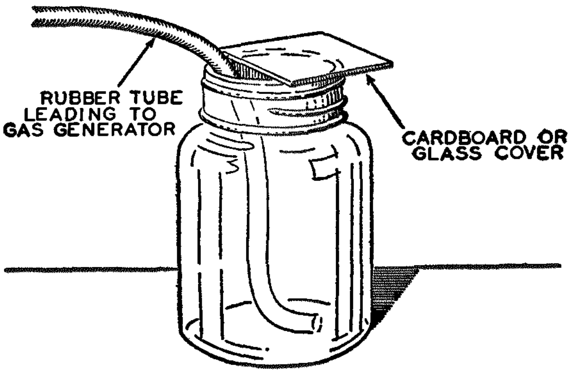
The illustration shows how the apparatus is set up. A round cake tin and some small jars or bottles in which to collect the gases are needed. Corked bottles are handy in some instances, as they can be filled with gas, then corked under water and kept until needed. The pan should be partly filled with water, just deep enough to seal the mouth of the jar when it is inverted in the pan.

In order to collect gas in a small jar, fill it with water. Then cover the mouth with a flat piece of cardboard, invert the jar, and set it upside down in the water in the pan. Slide the cardboard out from under without letting any water out or letting any air in. Place the end of the rubber tube leading from the gas generator under the edge of the inverted jar so that the gas can bubble up into the jar and displace the water.

To remove a gas-filled jar from the trough, remove the rubber tube and slide a piece of cardboard under the jar so as to close the mouth. Then lift the jar out of the pan. If it contains hydrogen, keep the jar in an inverted position. Otherwise it will escape. It is lighter than air.

Collecting Oxygen by Downward Displacement

Since oxygen is slightly heavier than air, you can fill a jar or bottle with it by downward displacement. The rubber tube through which the oxygen passes from the generator is inserted in the bottle or jar. The mouth of the jar is covered with a piece of cardboard. The oxygen settles to the bottom and pushes the air out at the top. Soon the bottle or jar is filled with oxygen. The only difficulty with this method is that you can not tell when the jar is full.



COLLECTING OXYGEN BY DOWNWARD DISPLACEMENT

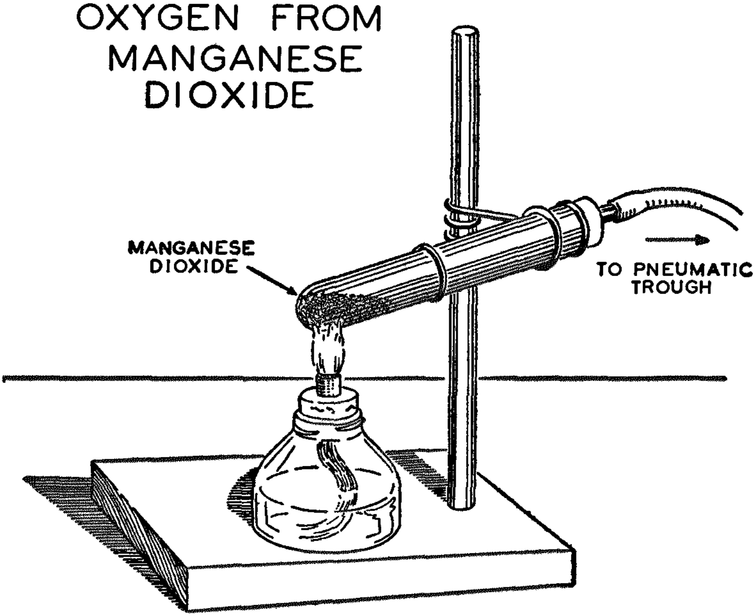
As a gas, oxygen resembles air in being colorless, odorless and tasteless. It is slightly heavier than air and consequently may be collected by downward displacement. If oxygen is led slowly into the bottom of an open jar, it will collect in a layer at the bottom. As its volume increases, it will push the air above it up and out of the jar. Hydrogen being lighter than air can not be collected in the same manner. It can be collected by *upward* displacement in an *inverted* jar. The lighter hydrogen pushes the heavier oxygen out at the *bottom.*

To Make Oxygen from Manganese Dioxide

MANGANESE DIOXIDE

Put 2 teaspoonfuls of manganese dioxide in a hard-glass test-tube fitted with a cork which is, in turn, fitted with a glass tube over which a rubber tube can be slipped.

Heat the test-tube until the manganese dioxide is red-hot and collect the gas which is evolved in a small bottle or jar in



Black oxide of manganese is a compound found in nature called pyrolusite by mineralogists and manganese dioxide by chemists. When heated, it loses part of its oxygen. It is another compound of manganese and oxygen containing the elements in different proportions. The illustration shows how to arrange a hard-glass test-tube on a ring stand for preparing oxygen from manganese dioxide.

a pneumatic trough. Test the gas with a glowing splinter. It is OXYGEN and the splinter will burst into flame.

When heated, manganese dioxide gives off part of its oxygen and changes to manganese sesquioxide and then to mangano-manganic oxide.

To Make Oxygen from Potassium Chlorate

POTASSIUM CHLORATE

Put 1 teaspoonful of potassium chlorate in a test-tube fitted with a stopper which is, in turn, fitted with a tube so that the gas which is generated in the test-tube may be led to a pneumatic trough.

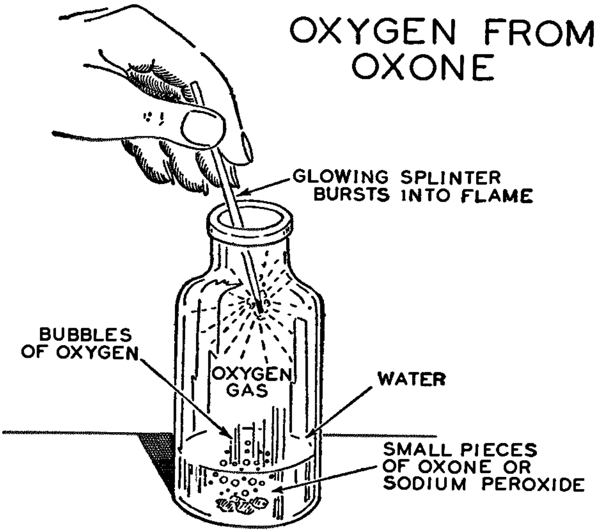
Heat the potassium chlorate gently. At first it will melt, forming a clear liquid. Then it will boil and a gas will be given off. Collect some of the gas in a small jar or bottle.

After gas has been produced for a time, the mass in the test-tube becomes solid. If at this point the heat is raised, gas will continue to be produced. In this second stage of the reaction, a larger amount of oxygen will be produced than in the first.

The point will finally be reached when the mass in the test-tube will remain unchanged, no matter how long heat is applied.

Heating potassium chlorate decomposes it. In the first stage of decomposition, POTASSIUM PERCHLORATE and OXYGEN are formed. In the second stage, the potassium perchlorate breaks up into POTASSIUM CHLORIDE and OXYGEN. The residue in the tube is POTASSIUM CHLORIDE.

Light a splinter of wood. Blow the flame out and thrust the still smoldering end into one of the jars or bottles which have been filled with the gas liberated by heating the potassium chlorate. It will burst into flame, proving that the gas is oxygen. A substance which is smoldering in air has its rate of combustion so speeded up that it bursts into flame in an atmosphere of nearly pure oxygen.



Oxone consists essentially of sodium peroxide. It gives off oxygen when treated with water. This is the simplest way to prepare oxygen.

Oxygen from Oxone

OXONE

Oxone is a fused form of sodium peroxide. In moist air or in the presence of water, it decomposes, forming sodium hydroxide and oxygen. Oxone is sold in cans. It is used in chemical analysis as an oxidizing agent, in bleaching, and for generating oxygen in hospitals, submarines, and mine-rescue apparatus.

Put a piece of Oxone the size of an ordinary marble in about ½ cup of water contained in a tall, narrow (pickle) bottle. Gas will be given off and fill the bottle. It is oxygen. Thrust a glowing splinter into the bottle of oxygen. The splinter will burst into flame.

Do not handle Oxone with the fingers or any combustible article. Use a metal spoon. Keep the can of Oxone tightly closed when not in use.

An Oxygen Generator

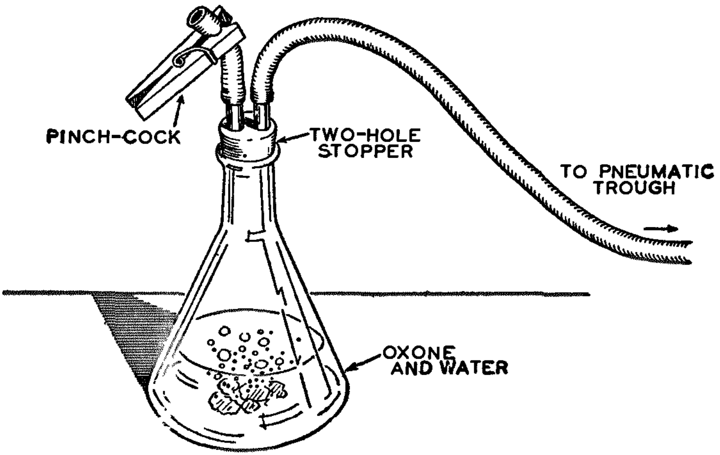
OXONE OR SODIUM PEROXIDE

A copious supply of oxygen may be secured by the reaction between Oxone or sodium peroxide and water in a homemade "generator."

The "generator" consists of an Erlenmeyer flask fitted with a two-hole rubber stopper. A short piece of glass tubing fits snugly into each hole. A piece of rubber tubing slipped over one of the glass tubes leads to a pneumatic trough. The other tube is fitted with a short length of rubber tubing which may be closed by pinching or with a wooden plug so as to keep the oxygen from escaping.

The Oxone or sodium peroxide is placed in the flask, and water is added as needed from time to time with a medicine dropper. It is not necessary to remove the stopper from the flask. The water may be dropped into the flask through the glass tube closed by the short piece of rubber tube.

AN OXYGEN GENERATOR



The short rubber tube, closed with a pinch-cock is used to introduce water into the generator. Oxone is placed in the flask and water added through the tube a few drops at a time.

Oxygen from Potassium Chlorate and Manganese Dioxide

MANGANESE DIOXIDE POTASSIUM CHLORATE

The method most commonly employed in making oxygen in the laboratory consists of heating a mixture of equal parts by weight of potassium chlorate and manganese dioxide.

Potassium chlorate contains by weight about 39 per cent oxygen, and by heating at 750 degrees Fahrenheit, or higher, you can drive out ALL the oxygen from the original material.

Manganese dioxide contains by weight about 36 per cent oxygen and by heating between 1,300 and 1,500 degrees

Fahrenheit, you can separate part of its oxygen. You *can not separate ALL the oxygen at any temperature.*

However, if you heat a MIXTURE of potassium chlorate and manganese dioxide, you get a very good flow of oxygen at 400 degrees Fahrenheit. Regardless of the amount of manganese dioxide used, ALL the oxygen comes from the potassium chlorate. The manganese dioxide does not take part in the decomposition but remains behind in its original form.

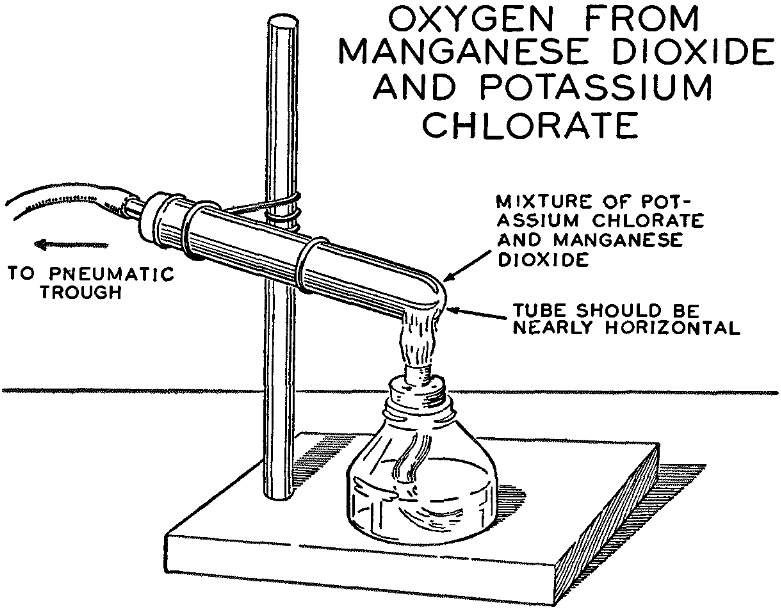
The manganese dioxide causes the potassium chlorate to give up its oxygen at a lower temperature than it does when heated alone. This occurrence illustrates a very fundamental chemical phenomenon called CATALYSIS. We call the material which causes the effect a CATALYTIC AGENT or CATALYZER.

In a *clean* dish, mix 2½ level teaspoonfuls of potassium chlorate with 1 level teaspoonful of powdered manganese dioxide. Put this mixture in a *clean, dry,* large test-tube supported from a stand and fitted with a stopper and tube arranged for the collection of gas in a pneumatic trough.

Support the test-tube as shown in the illustration. Notice that it is not quite horizontal, that the open end is slightly lower than the closed end. There is a good reason for this. It will prevent the moisture which condenses in the cold part of the tube from running back into the hot portion and cracking it.

Spread the mixture over the bottom half of the tube and heat very GENTLY. If there is any appreciable amount of dust or dirt in the mixture, the tube may BURST. Do not stand close to the tube when heating it. Hold the Bunsen burner at arm's length and wear a pair of goggles. Watch closely for sparks in the mixture. If it sparks, stop heating, throw it away, and make a new mixture, using more care to keep it free from dust and dirt.

When you start heating the mixture, the end of the delivery tube should be below the surface of the water in the pneumatic



Heating a mixture of potassium chlorate and manganese dioxide is the method of making oxygen most commonly employed in the laboratory. Equal parts by weight of the ingredients are used. The oxygen produced comes from the potassium chlorate. The manganese dioxide remains behind in its original form. Its presence causes the potassium chlorate to give up its oxygen at a lower temperature than if heated alone.

trough. The bubbles which escape for the first minute or so are expanded AIR and should not be collected.

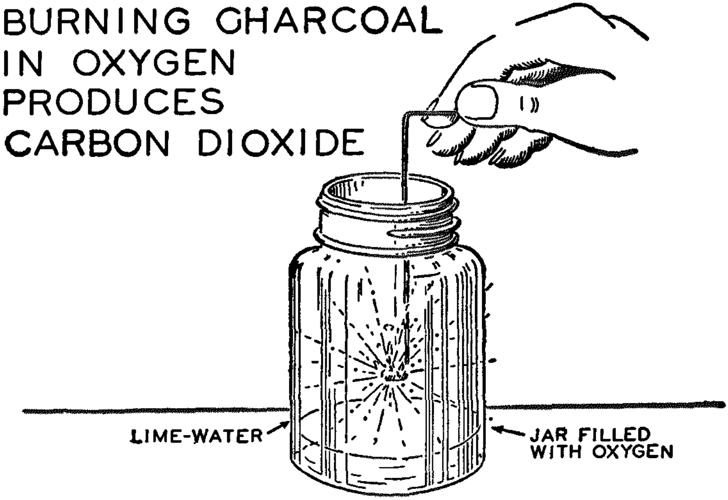
The flow of oxygen from the delivery tube will depend upon the amount of heat applied to the mixture in the test-tube. If the bubbles escape too fast, withdraw the heat. When you STOP heating, remove the delivery tube from the water, or water will be sucked up into the test-tube.

Fill several one-pint fruit jars with oxygen. Cover them and use them to perform the following experiments.

Burning Carbon in Oxygen

CHARCOAL OXYGEN

Bend a wire around a small piece of charcoal to form a handle. Ignite the charcoal in the flame of an alcohol lamp or



Oxygen does not act upon charcoal at ordinary temperatures but if a piece of red-hot charcoal is introduced into a jar filled with oxygen, the charcoal will burn with great brilliance. The product of burning charcoal in oxygen is carbon dioxide and its presence can be shown by a little lime-water in the bottom of the jar. If shaken so as to absorb some of the carbon dioxide, the lime-water will change from a clear solution to a cloudy mixture of water and calcium carbonate (chalk).

Bunsen burner. It will glow but will not burn with any particular vigor. While the charcoal is still glowing, thrust it into a jar of oxygen. Brilliant combustion will take place and continue until all the charcoal or all the oxygen is consumed.

When carbon burns in oxygen, it forms CARBON DIOXIDE. You can prove that the gas remaining in the jar is carbon dioxide by pouring in a little clear lime-water and shaking it up. The lime-water will become cloudy due to the formation of chalk or CARBONATE OF LIME. Since no other gas has the same effect upon lime-water, this is visible proof of the presence of CARBON DIOXIDE.

Sulfur Burns Brilliantly in Oxygen

SULFUR OXYGEN

Make a "deflagrating spoon" by bending one end of a narrow strip of sheet-iron (cut from a tin can) at a right angle. The strip should be approximately 8 inches long and 3/8 inch wide.

Place some sulfur on the end of the spoon and ignite it. It will melt first and then take fire. Notice whether it burns with ease or difficulty. Observe the odor of the fumes which are given off.

Before all the sulfur has burned off the spoon, thrust it into a jar filled with oxygen. It will be seen that sulfur burns much more readily in the oxygen than in the air. It will burn with a brilliant blue flame.

When the combustion has ceased, smell the contents of the jar very cautiously. The odor is that of SULFUR DIOXIDE. When sulfur burns in air or oxygen it forms sulfur dioxide.

Steel Wool Burns in Oxygen

STEEL WOOL OXYGEN

Bend a piece of iron wire at one end into a small hook. Fasten a small wad of fine steel wool on the hook and hold it in a flame until it is red-hot. Before it has a chance to cool, plunge the red-hot wool into a jar of oxygen. It will burn brilliantly.

White-hot drops of molten steel which drop to the bottom of the jar are apt to crack the glass. Before introducing the red-hot steel wool to the jar, throw in enough sand to cover the bottom. This is merely a mechanical means of protection against the hot products of the combustion.

Notice the reddish-brown substance on the end of the wire or on the sand in the jar. There may be a fine reddish-brown powder deposited on the sides of the jar. This is OXIDE OF IRON.

Potassium Nitrate Is an Oxidizing Agent

IRON FILINGS POTASSIUM NITRATE

Potassium nitrate, or saltpeter, is a powerful oxidizing agent. It is an ingredient of gun-powder and of most fireworks, its purpose in such mixtures being to furnish oxygen.

To illustrate the oxidizing power of potassium nitrate, mix 2 teaspoonfuls of that substance with ½ teaspoonful of iron filings. Place the mixture in a hard-glass test-tube and heat it in the flame of a Bunsen burner.

First, the potassium nitrate will melt in its own water of crystallization. Then a brisk combustion will take place. The iron will be oxidized by the oxygen of the potassium nitrate and will be changed into IRON OXIDE.

Sulfur Burns in Potassium Nitrate

POTASSIUM NITRATE SULFUR

Put a heaping teaspoonful of potassium nitrate in a hard-glass test-tube. Fasten the tube in a clamp attached to a ring stand. Do not hold it in the hands or with a test-tube holder. Considerable heat is generated by this experiment.

Heat gently over a Bunsen burner or alcohol lamp and when the potassium nitrate melts, remove the flame and quickly drop in a small piece of sulfur. Be prepared for a very brilliant flash; the sulfur will burn with an intensely brilliant flame.

A pan of sand or a sheet of tin or of asbestos should be placed under the tube before the sulfur is dropped in. The heat developed during this reaction is often sufficient to melt a soft-glass test-tube.

Charcoal Will Burn in Potassium Nitrate

POTASSIUM NITRATE CHARCOAL

Potassium nitrate is so capable of furnishing oxygen, and charcoal is so anxious to acquire it, that they will burn if heated together.

Melt two teaspoonfuls of potassium nitrate in an evaporating dish or porcelain crucible. Considerable heat will be required. Drop in some small bits of charcoal. They will take fire and burn vigorously.

Sugar Will Burn in Potassium Chlorate

An OUTDOOR Experiment

POTASSIUM CHLORATE SUGAR SULFURIC ACID

Sugar belongs to that class of chemicals called carbohydrates. Sugar will burn, and if furnished with sufficient oxygen, will burn with a flash. Potassium chlorate, if mixed with sugar, will furnish sufficient oxygen to burn the sugar very rapidly.

Put ½ teaspoonful of potassium chlorate on a sheet of paper and, using a strip of cardboard as a spoon, gently mix it with ½ teaspoonful of powdered sugar.

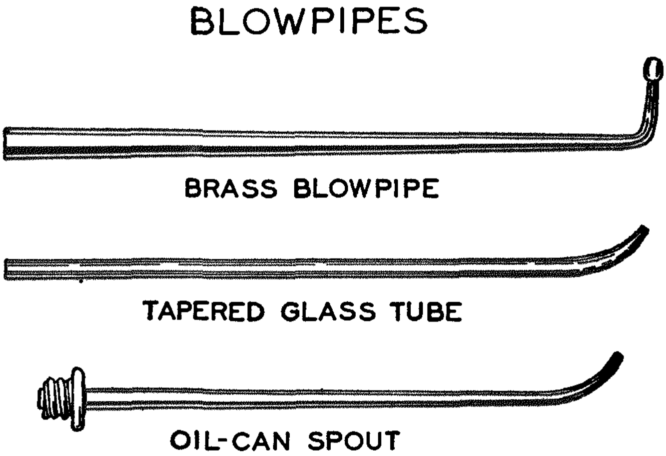
Take the mixture of chlorate and sugar outdoors and pour it in a small heap on a flat stone. Then, standing back so that your face is not over the mixture, let three or four drops of sulfuric acid fall from the end of a glass rod onto the mixture. A lively combustion, tinged with the violet-colored flame characteristic of potassium, will take place.

MAKING AND USING A BLOWPIPE

A blowpipe is a tube through which air is blown into a flame by means of the mouth. The composition of a substance can often be determined by heating it in a blowpipe flame and noticing its conduct. Some metals, when placed upon a block of charcoal and heated in a blowpipe flame, form characteristic films of oxides on the charcoal. It is possible to determine the presence of certain substances from the color of the film of oxide. The blowpipe is therefore sometimes of great value in chemical analysis.

A blowpipe is usually constructed of metal in the shape shown at the top of the illustration below. A blowpipe satisfactory for many purposes may be made from glass tubing by drawing one end down to a smaller size.

You can purchase an oil-can with a long curved spout at the dime store for 10 cents. Procure one of these cans, remove the spout, and use it as a blowpipe.



At the top of the illustration is shown an inexpensive form of blowpipe which can be purchased from a laboratory supply-house. Below are home-made forms. The glass tube is curved and tapered at one end. It is rather a poor substitute for a metal blowpipe. Unless used with skill, the end may soften in the flame and close. A long oil-can spout makes a good blowpipe. If the hole at the small end is too large, close it slightly by squeezing or hammering.

A blowpipe may be used with the flame of a Bunsen burner, candle, or alcohol lamp. When a Bunsen burner is used, cut off the supply of air which usually enters the holes at the base. The flow of gas should be regulated so that the flame is about 2 inches high.

Place the tip of the blowpipe on the edge of the Bunsen burner slightly in the base of the flame. Blow gently into the blowpipe and the flame will be forced into a long needle of fire. There are two main cones in a blowpipe flame. The extreme tip of the outer cone is the most efficient oxidizing flame. The tip of the inner cone is the most efficient reducing flame. The different effect of the two parts of the flame will be shown in some of the experiments which follow.

When using a blowpipe, puff out your cheeks, using them as a sort of bellows. The air should be forced out by the action of the cheeks rather than by the pressure of the lungs. A little practice will make it possible to blow for a long time and breathe naturally while doing it.

Oxidizing Zinc with a Blowpipe Flame

ZINC CHARCOAL

Lay a small piece of zinc in a hollow on a block of charcoal and heat it for a few minutes with the OXIDIZING blowpipe flame. First the zinc will melt. Then it will start to oxidize. Notice the bluish color it gives to the blowpipe flame. The smoke and the film which form on the charcoal are ZINC OXIDE. Notice also the color of the zinc oxide when it is hot and when cold. It is yellow when hot and white when cold, in the same way that mercuric oxide is brown while hot and red when cold.

The principal use for zinc oxide is in making paint, zinc white or Chinese white. Zinc paint does not become darkened by hydrogen sulfide in the atmosphere as does lead paint. The zinc oxide used in paint is prepared by burning zinc.

A paste made of zinc oxide and zinc chloride is sometimes employed by dentists for filling teeth.

Oxidizing Lead with a Blowpipe Flame

LEAD CHARCOAL

Heat a small piece of metallic lead on a block of charcoal, using the oxidizing flame of the blowpipe. When the lead has melted, notice the formation of yellow lead monoxide which forms a coating or film on the charcoal in the neighborhood of the metal.

There are five different oxides of lead. The yellow monoxide or litharge is formed by heating lead in the air at a low temperature. A red oxide called minium or red lead is formed at high temperatures.

Reducing Lead Oxide with Blowpipe Flame

SODIUM CARBONATE LEAD MONOXIDE (LITHARGE)

Select a block of charcoal the upper surface of which is flat and level when the block is resting on the table. Make a hollow spot in the center of the surface by pressing a penny against it and twisting it around a few times.

Mix thoroughly 1 level teaspoonful of sodium carbonate with 1 level teaspoonful of litharge or lead monoxide on a sheet of paper. Add a few drops of water, just enough to make the mixture stick together and prevent it from being blown away when you first direct the blowpipe against it. Fill the cavity in the charcoal with the mixture and heat it in the reducing flame produced by the blowpipe. Continue to use the reducing flame until several bright silvery globules appear in the cavity. Then stop the heating.

It should be explained at this point that the sodium carbonate in the mixture does not take any *chemical* part in the reaction. Its purpose is purely *physical.* It serves as a FLUX. It melts easily and holds the powdered litharge in position so that the flame may play upon it.

After the mixture has cooled, scrape it out of the cavity and break it up. Treat it with a little water so as to dissolve as much as possible. Then pick out the beads. Rub one across a sheet of white paper and observe the mark it makes. Hit one of the beads with a hammer. It is not hard. The beads are lead. They have been produced by REDUCING lead oxide. The reducing flame of the blowpipe took the oxygen out of the lead oxide and used it to burn some of the partially consumed gas in the inner cone of the flame. Metallic lead was left behind.

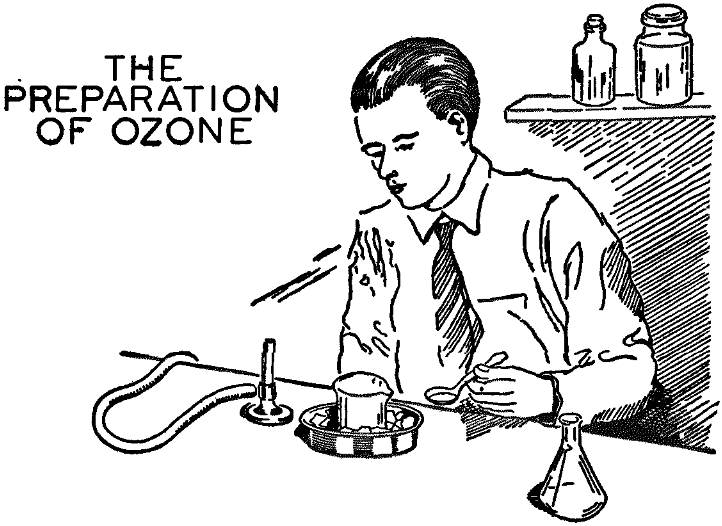
Assayers frequently test ores by the blowpipe method. By heating the ore in a blowpipe flame, they obtain the metal contained in the ore.

Reducing Lead Acetate

LEAD ACETATE CHARCOAL

Moisten a heaping teaspoonful of lead acetate with a few drops of water (just enough so that it will hold together and not be scattered by the blast from a blowpipe flame).

Put the moistened lead acetate into a small hollow on a piece of charcoal and heat it with a blowpipe. The acetate will melt, then become brown. Finally it will char and a moment later globules of metallic lead will appear on the charcoal. The acetate has been reduced.



An interesting thing about oxygen is that it can exist in the second form called ozone. Ozone is one and one-half times as heavy as ordinary oxygen and behaves differently. For example, silver is not affected by oxygen but is rapidly oxidized by ozone. Ordinary oxygen is the more stable form. Ozone soon changes to oxygen at ordinary temperatures and does not remain in the air long after its formation.

Preparation of Ozone

BARIUM PEROXIDE

SULFURIC ACID

Pour 1 teaspoonful of concentrated sulfuric acid into a small beaker. Place the beaker in a dish and pack some cracked ice around it. Slowly and carefully shake 1 teaspoonful of barium peroxide into the beaker. The reaction of the barium peroxide with the cold sulfuric acid produces OZONE, a form of oxygen which has a strong, unpleasant smell. If the mixture is not thoroughly cold, oxygen, rather than ozone, will be formed. You will be able to recognize the characteristic odor of ozone in the air immediately above the beaker. The smell might be likened to a mixture of sulfur dioxide, garlic, and chlorine.

CHAPTER V

EXPERIMENTS WITH GASES AND SOME OF THEIR COMPOUNDS

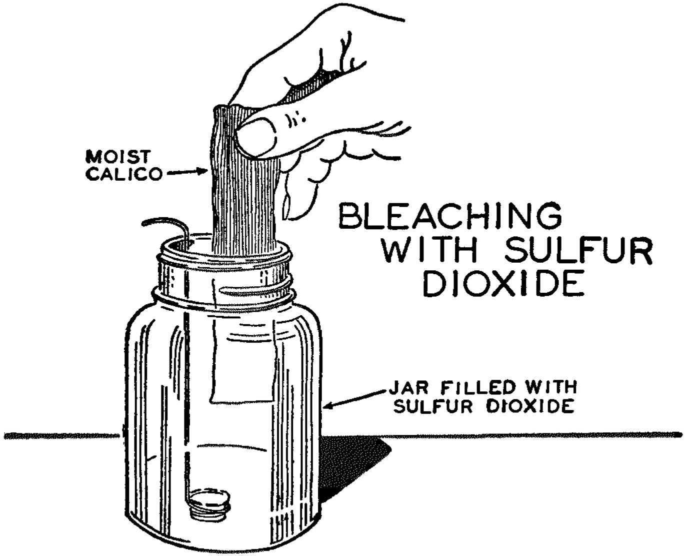
MANY of the chemists' most prized possessions are made of gases. Sulfuric, nitric, and hydrochloric acids are compounds of gases. Ammonia, a gas which consists of nitrogen and hydrogen, is the starting point of all ammonia compounds. Sulfur dioxide is one of the raw materials for making sulfuric acid. Sulfuric acid gives the chemist his long list of sulfates and is an agent for making many other chemicals.

The manufacture of nitrogen, oxygen, chlorine, helium, acetylene, carbon dioxide, and hydrogen is a large industry. Of course, nitrogen, oxygen, hydrogen, helium, and chlorine are elements and, as such, are not actually manufactured. They are merely purified and compressed.

That gases are important and fill a large place in the industrial scheme of things is apparent from the fact that many *billion pounds* of gases are produced and sold in the United States every year.

Bleaching with Sulfur Dioxide

When sulfur burns in oxygen or air, sulfur dioxide is formed. *Pure* sulfur dioxide is a colorless gas. When sulfur burns in air, pure sulfur dioxide is not produced. The nitrogen in the air seems to favor the production of a small amount of sulfur trioxide as well. The presence of sulfur trioxide in sulfur dioxide accounts for the slightly foggy appearance of the gas.



Sulfur dioxide is a colorless gas of unpleasant odor. It is a good bleaching agent and is used for bleaching paper, wool, silk, straw, etc In some instances the bleaching action is quite complicated but in others it is simply due to the fact that the sulfur dioxide withdraws oxygen from the colored substances with the consequence that they become colorless products.

Paper, pulp, straw, silk, wool, and other materials which would be injured by chlorine can be bleached safely with sulfur dioxide.

Put some sulfur into a deflagrating spoon made by twisting a piece of wire around a metal bottle cap. Ignite the sulfur and place it in a large, wide-mouthed bottle or a fruit jar. The burning sulfur will fill the jar with SULFUR DIOXIDE.

Place a *wet* red carnation, rose, green leaf, blue flower, or a piece of wet, colored calico in the sulfur dioxide jar. In order for the sulfur dioxide to act, the article must be wet. The flower, cloth, or leaf will be bleached white.

The color of materials bleached by sulfur dioxide often returns on exposure to air. The color of bleached rose-leaves gradually returns when the rose-leaves are exposed to the air

The Preparation of Chlorine

BLEACHING POWDER	DILUTE SULFURIC ACID

Chlorine is about two and a half times as heavy as air. It can be prepared in a test-tube. As the chlorine forms in the bottom of the test-tube, it will push the air out at the top. Chlorine is heavy enough to be poured from the tube.

Put a saltspoonful of bleaching powder in the bottom of a test-tube. If you use more than this small amount, you will make too much chlorine. Set the tube in a rack and add a small amount of storage-battery electrolyte (dilute sulfuric acid).

As the chlorine begins to fill the tube, test from time to time with a strip of moistened blue litmus-paper. When the litmus-paper, held in the mouth of the tube, turns white, there is enough chlorine in the tube to try an experiment in bleaching a piece of cloth.

Chlorine is an irritating greenish-yellow gas which is not dangerous to experiment with if you do not use larger quantities of chlorine-producing chemicals than the instructions in this book call for and if you do your experimenting in a room with a window open.

The Bleaching Action of Chlorine

BLEACHING POWDER DILUTE SULFURIC ACID

Hang a narrow strip of dry, colored calico in a test-tube containing chlorine. Little or no bleaching occurs. If either the calico or the chlorine is moist, however, the color of the calico will disappear.

If no bleaching of the calico strip occurs, lift it out of the tube and wet it thoroughly with water. When it is again immersed in the chlorine gas, the color will begin to fade. After a while, the colored calico will be white or light yellow.

Chlorine itself does not actually bleach. It causes bleaching to occur. The presence of water is necessary. The chlorine reacts with the water and releases nascent or atomic oxygen. It is really the action of the nascent oxygen on the dye in the cloth that does the bleaching. The oxygen reacts with the dyes and forms colorless compounds.

Nitric Oxide

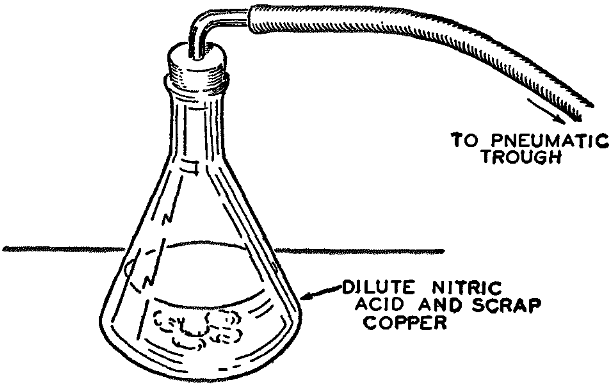
COPPER NITRIC ACID

There are five oxides of nitrogen: nitrous oxide, nitric oxide, nitrogen trioxide, nitrogen peroxide, and nitrogen pentoxide.

Nitric oxide may be prepared by the action of dilute nitric acid on copper. The gas which is generated is not pure nitric oxide; it contains some nitrous oxide and some nitrogen.

Put some small pieces of copper wire, sheet-copper, copper turnings or filings into a small Erlenmeyer flask. Add 5 or 6 teaspoonfuls of water and then pour nitric acid in *very slowly.* When enough acid has been added, a brown gas will be evolved. If too much acid is added, the liquid may froth and be forced out of the bottle.

NITRIC OXIDE



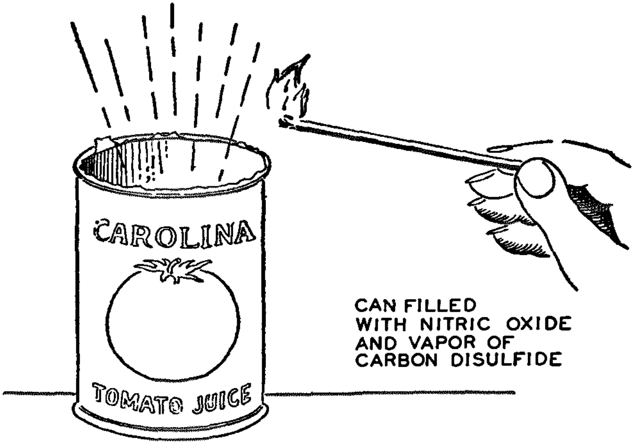
Nitric oxide is a colorless, transparent unstable gas whose most remarkable property is its ability to combine directly with oxygen when the two are brought together. The product is a poisonous reddish-brown gas called nitrogen peroxide. Nitric oxide does not burn and its instantaneous change into nitrogen peroxide when given the opportunity is not accompanied by the evolution of light although heat is evolved.

The first nitric oxide produced in the flask by the action of nitric acid and copper will mix with air in the flask and immediately change into reddish nitrogen peroxide. When all the oxygen in the air in the flask has been consumed, the nitric oxide produced will remain unchanged.

Close the flask with a one-hole stopper fitted with a tube so that the gas can be collected over water in a pneumatic trough. At first the flask will be filled with the reddish-brown gas, but soon the gas generated will be colorless.

Collect, over water, two pint jars of nitric oxide. Put a burning splinter of wood into one of the jars. The splinter will be extinguished. Nitric oxide does not burn, and most burning substances are extinguished when introduced into it. It seems strange that *nitrous* oxide will support combustion and nitric oxide, containing twice as much oxygen in proportion to nitrogen, will not. But such is the case.

AN EXPERIMENT WITH NITRIC OXIDE



Carbon disulfide is one of the few substances which will burn in nitric oxide.

An Experiment with Nitric Oxide

NITRIC OXIDE CARBON DISULFIDE

Phosphorus is one of the few substances which will burn in *nitric oxide* but phosphorus is too dangerous for the young experimenter to have around.

You can perform a surprising experiment with nitric oxide and a few drops of carbon disulfide.

In a jar of nitric oxide place a few drops of carbon disulfide. Close the jar for a few minutes to allow the carbon disulfide to evaporate and its vapor to mix with the nitric oxide.

Bring a burning splinter of wood or a taper to the open mouth of the jar. There will be a flash and the mixture will burn with a brilliant light. Try this experiment in a darkened room if possible.

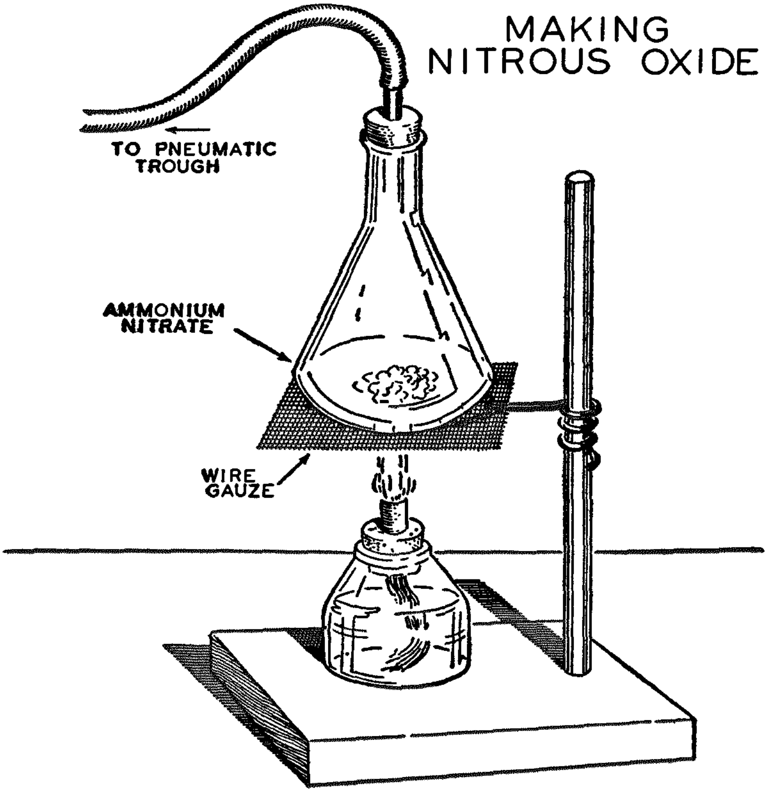
Carbon disulfide is very inflammable and evaporates rapidly. It should never be poured out of its container if an open flame is in the same room.

Nitrous Oxide

AMMONIUM NITRATE

Put 4 teaspoonfuls of ammonium nitrate into an Erlenmeyer flask provided with a tight-fitting one-hole stopper through which a short length of glass tubing passes. A piece of rubber tube attached to the glass tube in the cork leads to a pneumatic trough. The gas which is generated in the Erlenmeyer flask is to be collected in a pint jar over hot water in the same manner as the oxygen generated by heating potassium chlorate and manganese dioxide (see page 91). The water in the pneumatic trough must be *hot*. This is important. Otherwise, much of the gas will dissolve in the water.

Set the flask containing the ammonium nitrate on a double thickness of wire gauze on a ring stand. Heat the bottom of the flask *gently*. The ammonium nitrate will melt and small bubbles of gas will begin to escape. At this point the heat



Nitrous oxide is produced by heating ammonium nitrate or a mixture of a salt of ammonium nitrate.

Nitrous oxide is a colorless gas having a slightly sweetish taste. It is somewhat soluble in water so that when collected over water there is always considerable loss. A glowing splinter of wood bursts into flame when introduced into nitrous oxide. Other combustibles burn in it much the same as in oxygen. However, metals do not rust in nitrous oxide and it can not be utilized as a source of oxygen by the hemoglobin of the blood.

must be carefully controlled so that the evolution of gas is not too rapid. If you heat ammonium nitrate too rapidly, it may explode. This does not happen until a temperature of

480 degrees Fahrenheit is reached. It is perfectly safe to heat ammonium nitrate *gently.*

When the jar has been filled or when the generation of gas ceases, remove the heat from under the flask.

Remove the jar from the pneumatic trough and insert a glowing splinter of wood into the gas. It will burst into flame and burn with almost as much brilliance and energy as if it were in oxygen. Although the gas contains oxygen, it is not oxygen. It is NITROUS OXIDE.

Although nitrous oxide will support combustion, metals do not rust in it. The red blood corpuscles of the blood are unable to use it as a source of oxygen. If nitrous oxide without any air or oxygen is taken into the lungs, insensibility follows, since nitrous oxide furnishes no oxygen. By mixing it with a sufficient amount of air to sustain life, it is employed as an anesthetic for minor operations and for extracting teeth.

Burning Steel Wool in Nitrous Oxide

NITROUS OXIDE SULFUR STEEL WOOL

Fill a wide-mouthed bottle with nitrous oxide by heating ammonium nitrate as described in the previous experiment. Keep the bottle tightly corked until you are ready to insert the hot steel wool.

Push a stiff piece of wire having a hook on the end into a cork of the right size to fit the bottle of nitrous oxide. The hook should be a little more than half-way to the bottom when the cork is put into the mouth of the bottle.

Arrange a small wad of steel wool on the hook on the end of the wire and heat it red-hot. Quickly dip the wool into a little powdered sulfur in such a way that a small quantity of burning sulfur will cling to the wool and keep it hot on its way to the bottle of nitrous oxide.

Quickly thrust the *hot* steel wool into the bottle of nitrous oxide. The steel will burn with great brilliance and a shower of sparks, much as it burns in oxygen.

When all combustion has ceased, remove the cork and the wire and cautiously test the odor in the bottle. Do you recognize the odor of SULFUR DIOXIDE? There may be a trace of it in the jar.

Making Illuminating Gas

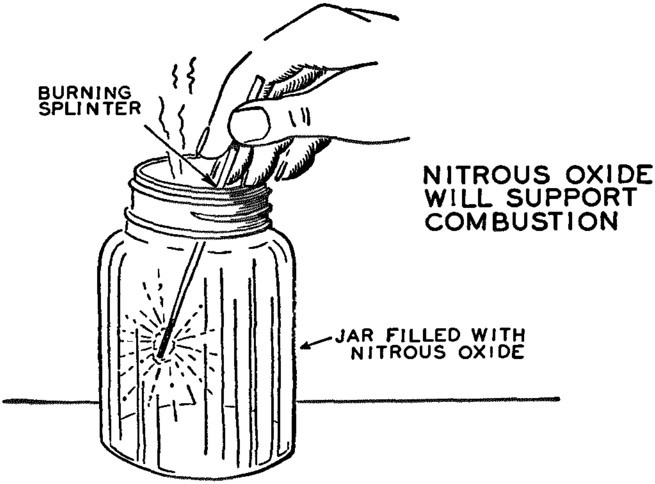
SOFT COAL

Of all the forms of carbon, coal is by far the most important. The two most used kinds of coal are *anthracite* and *bituminous*. Both were formed long ago from the vegetation which once covered the regions in which coal is now found. It is thought that the vegetation became covered with earth deposits which made it impossible for the vegetation to secure oxygen, and so it decayed slowly. Under the influence of moisture, pressure, and temperature, the dark solid which we know as coal was formed.

Anthracite coal is hard and contains very little gas. In bituminous or soft coal, a large amount of gas exists.

Put 2 teaspoonfuls of powdered soft coal in a hard-glass test-tube. Fit the test-tube with a stopper and a delivery-tube leading to a pneumatic trough. Support the tube so that the bottom can be heated.

Use a Bunsen burner or an alcohol lamp to heat the powdered coal in the tube and collect one or two jars of the gas which is evolved. The gas in the jars will burn if ignited with a match. The gas issuing from the end of the delivery tube will burn with a yellow flame. This gas is a mixture of several compounds of carbon and hydrogen. It is ILLUMINATING GAS.



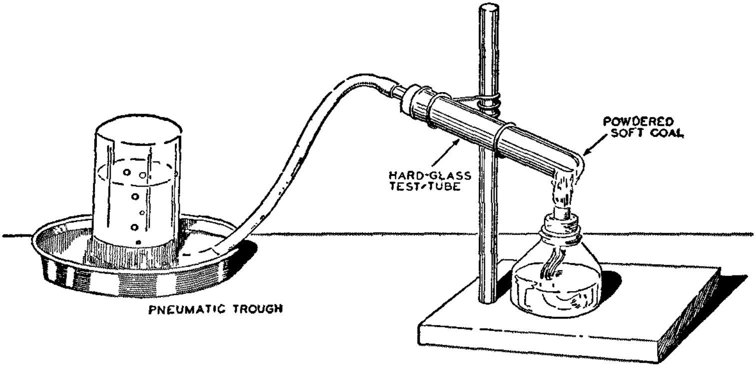
A glowing splinter thrust into a jar of nitrous oxide will burst into flame and burn almost as violently as if the jar contained pure oxygen. Although nitrous oxide will support combustion it will not support respiration.

As soon as the coal in the test-tube ceases to give off gas, take away the flame, remove the pneumatic trough, and when the tube has cooled, break it open and examine the contents. It is COKE.

Besides carbon, hydrogen, oxygen, and nitrogen, coal contains a certain quantity of mineral substances. If heated or burned in the open air, bituminous coal is almost entirely consumed. But it leaves its minerals behind in the form of ashes.

But when heated out of contact with air, as in this experiment, it is subjected to DESTRUCTIVE DISTILLATION and the results are entirely different. The products are various gases and volatile substances, an interesting tarry liquid, and coke.

MAKING ILLUMINATING GAS



To-day, "illuminating" gas is used almost entirely for fuel. There are two methods of making gas from coal. The coal gas method is to heat soft coal in retorts from which the air is shut out, as in the experiment illustrated above. Coal gas is chiefly a mixture of methane and hydrogen. Important by-products are coke, coal-tar, ammonia and liquid hydrocarbons.

Water gas is produced by blowing steam through white hot coal. In some large gas works, both coal gas and water gas are made and mixed to form illuminating gas.

Illuminating gas is manufactured by the destructive distillation of soft coal. In addition to the illuminating gas, various chemicals such as ammonia, coal-tar, and coke are obtained.

From the thick, black, foul-smelling coal-tar the chemist can produce important dyes, drugs, explosives, naphthalene, benzene, and toluene.

Distilling Wood

SAWDUST

Arrange a hard-glass test-tube and a pneumatic trough in the same manner as in the previous experiment.

Put 2 or 3 heaping teaspoonfuls of dry sawdust in the test-tube and heat. Collect the gas which is given off in a small jar and test it to see whether or not it will burn. It will.

When the flow of gas has ceased, remove the delivery tube from the water and plug the end or compress the tube with a pinch-cock so that no air can enter the test-tube. When the tube and its contents have cooled, remove the cork and shake out the CHARCOAL which is contained in it.

Charcoal is made by heating wood in much the same manner that coke is made from coal, that is, without contact with air. The volatile products driven off from the wood are combustible gases, wood alcohol, acetic acid, acetone, and tar.

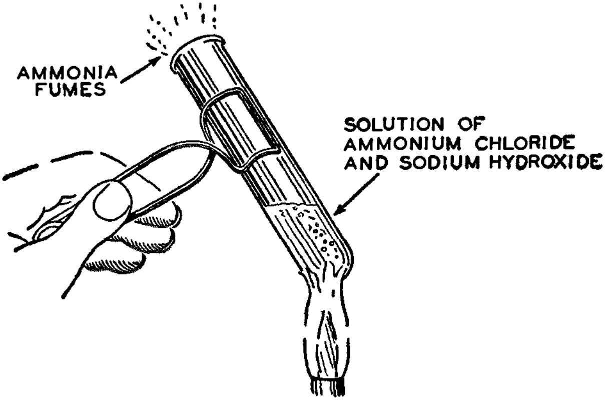
There is more than one variety of charcoal. So far, we have been talking about *wood* charcoal. Charcoals are used as fuel, for gun-powder, for carbonizing, and as an absorbent.

Gun-powder charcoal is usually made from alder or willow wood. Absorbent charcoals are usually made from cocoanut shells. Animal charcoal is made by heating bones in a closed retort.

How to Make Ammonia

AMMONIUM CHLORIDE CAUSTIC SODA

Vast quantities of ammonia are required for industrial purposes. Probably the most familiar form of the compound is the one used in the household for cleansing purposes. It is called household ammonia but it is actually an ammonia solution, called AMMONIUM HYDROXIDE by the chemist.



When sparks from an induction coil are passed through a mixture of hydrogen and nitrogen, small quantities of an extremely useful compound of nitrogen and hydrogen called ammonia are formed. Ammonia is a colorless gas with the pungent odor familiar in smelling salts. It is liberated in the putrefaction of animal matter.

Ammonia is produced commercially as a by-product in the manufacture of illuminating gas and coke. In the laboratory, it is made by heating a mixture of sodium hydroxide or slaked lime and some salt of ammonium such as ammonium chloride.

Ammonia itself is a gas. Although so-called household ammonia is a liquid, it is not to be confused with true liquid ammonia, which is made by compressing the gas.

Ammonia, compounded of nitrogen and oxygen, is found where the decomposition of animal and vegetable matter takes place. Consequently, traces of ammonia can be detected in the soil, in the air, in rain-water, and in the waters of rivers and lakes.

To a test-tube half-filled with water, add ½ teaspoonful of ammonium chloride (sal ammoniac). When this has dissolved, pour in a few drops of a solution of caustic soda (sodium hydroxide). Heat the mixture until it starts to boil. Ammonia can now be detected by its odor as it escapes from the mouth of the test-tube. The escaping gas will turn moistened red litmus-paper blue.

Hydrogen, an Ingredient of Ammonia

IRON FILINGS POTASSIUM HYDROXIDE

Mix 1 teaspoonful of iron filings with ½ teaspoonful of potassium hydroxide. Put this mixture in a test-tube fitted with a stopper and a delivery tube so that the gas which is to be generated can be collected over water in a pneumatic trough.

Heat the mixture contained in the test-tube and collect the gas which is evolved in a small bottle. When the bottle is filled, remove it from the pneumatic trough, taking care to keep it in an inverted position. The gas which it contains is lighter than air and will escape quickly if the mouth of the bottle is turned upward. Hold a lighted match to the mouth of the bottle and turn it upward. The gas will burn. It is HYDROGEN. It came from the potassium hydroxide. Some of the oxygen of the potassium hydroxide combined with the hot iron, thus releasing hydrogen.

Nitrogen, an Ingredient of Ammonia

IRON FILINGS POTASSIUM NITRATE

Mix 1 teaspoonful of iron filings and ½ teaspoonful of potassium nitrate. Put this mixture in a test-tube fitted with a delivery tube so that the gas which will be generated can be collected over water in a pneumatic trough.

Heat the mixture and collect the gas which is evolved in a small bottle. When the bottle is filled, turn it so that the mouth is up and insert a lighted match. The match will be extinguished. The gas in the tube is NITROGEN.

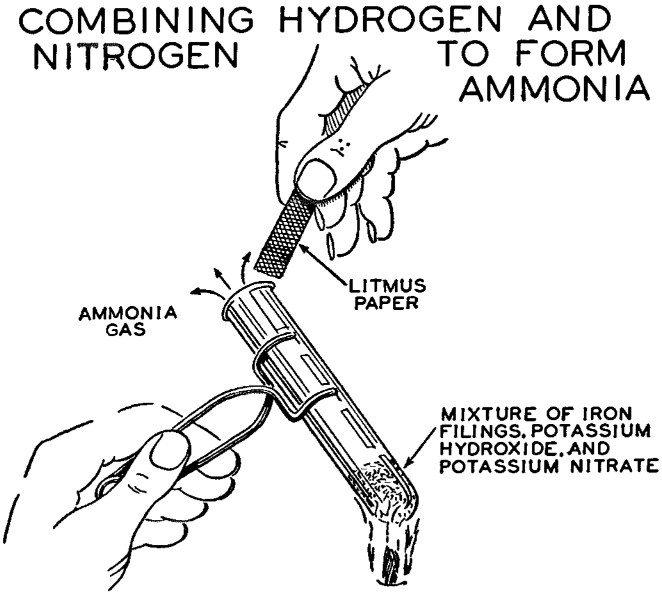
Combining Hydrogen and Nitrogen to form Ammonia

POTASSIUM HYDROXIDE IRON FILINGS
POTASSIUM NITRATE

By combining the reactions in the last two experiments and producing hydrogen and nitrogen simultaneously, ammonia results.

Mix 2 teaspoonfuls of iron filings with ½ teaspoonful of potassium hydroxide and ½ teaspoonful of potassium nitrate. Put this mixture in an open test-tube and heat.

The pungent odor of ammonia will issue from the test-tube. A strip of moistened red litmus-paper held over the test-tube will be turned blue. Hydrogen and nitrogen have combined to form ammonia. This union takes place only at the moment when the gases are liberated from another combination, or



The common laboratory method of preparing ammonia is to break down an ammonium salt and thereby release ammonia. Another method is to synthesize or combine hydrogen with nitrogen so as to form ammonia, as illustrated above. That ammonia is produced by this reaction is proved by the odor which issues from the test-tube. A piece of wet red litmus paper held in the mouth of the test-tube will change to blue. The ammonia gas dissolves in the water in the wet paper, forming a weak base and changing the color of the litmus.

in what is called a *nascent* state. If they do not meet until afterwards, no union takes place.

Ammonia Precipitates Hydroxides

HOUSEHOLD AMMONIA ALUM

Household ammonia, as already explained, is a solution of ammonia gas in water. Under ordinary conditions of temperature and pressure, water can dissolve 700 times its own volume of ammonia. Concentrated ammonium hydroxide is too strong to use for ordinary purposes and the household variety is an impure, dilute solution.

Ammonium hydroxide will precipitate hydroxides of many of the metals from their salts.

Dissolve a crystal of alum in a test-tube half filled with water. Add household ammonia. A gelatinous precipitate of ALUMINUM HYDROXIDE will be formed. Aluminum hydroxide is used in medicine and for purifying drinking water.

Aluminum hydroxide is formed by the addition of any hydroxide solution to an aqueous solution of an aluminum salt.

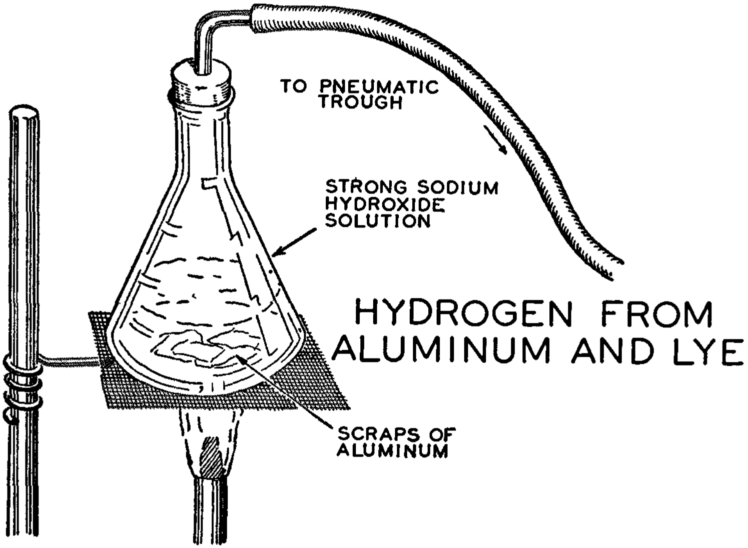
To Make Hydrogen from Aluminum, Water, and Lye

ALUMINUM SODIUM HYDROXIDE

Place some scraps of aluminum in a small Erlenmeyer flask and add a strong solution of sodium hydroxide in water (lye). Sufficient solution to cover the aluminum is all that is necessary. Fit the flask with a delivery tube so that the gas which is evolved can be collected over a pneumatic trough.

Apply a gentle heat to the flask to start the reaction. The heat may then be removed and the reaction will proceed without further assistance. HYDROGEN gas and SODIUM ALUMINATE will be formed.

The first bubbles which escape from the delivery tube will be largely air. The hydrogen itself will contain traces of sodium and will burn with a yellowish flame, not a colorless flame, as would be the case with sodium-free hydrogen. This slight impurity will make no difference when the hydrogen is used for ordinary experiments.



Hydrogen, the lightest of all gases, might be called the "business end" of all acids, and together with oxygen, the "business end" of bases. It is very widely used in chemical industry. The illustration above shows one of the many methods of preparing hydrogen.

Hydrogen from an Alkali and a Metal

SCRAP IRON SODIUM HYDROXIDE

Hydrogen can be generated by the interaction of a metal and an alkali. Sodium hydroxide and iron will produce hydrogen. You can use ordinary household lye and iron filings, iron tacks, or small nails.

Mix a heaping teaspoonful of lye with an equal amount of iron and place the mixture in a test-tube. Mount the test-tube in a clamp in a vertical position. Apply a gentle heat by moving a small flame around the sides and across the bottom of the tube. This gentle application of heat is intended to expel moisture from the lye. As soon as steam ceases to escape from the tube, increase the heat and the reaction will begin. Test the vapor in the tube from time to time by holding a lighted match to its mouth. SODIUM FERRATE and the gas HYDROGEN are formed by the reaction. At first the gas will burn only momentarily, but finally it will burn steadily. In a large test-tube the hydrogen will burn at the surface of the mixture of iron and sodium hydroxide, but if the tube is small, it will burn at the mouth.

CHAPTER VI

CHEMICAL TESTS

A CHEMIST is often confronted with a substance or a compound the exact nature of which he does not know but which he would like to ascertain. Since he is familiar with, or has books which describe, the manner in which various elements and compounds act when treated with other chemicals, he is usually able to make tests which reveal the composition of the unknown substance. The process of determining the composition of a substance is called *analysis.*

There are two varieties of chemical analysis. Determining the QUALITY of a substance or the nature of its elements is called *qualitative* analysis. Determining the QUANTITY or how much of any of the various elements there is in a compound is *quantitative* analysis.

Quantitative analysis is not a simple process. It requires training and experience. There are many highly specialized branches of quantitative analytical chemistry. In fact, you can not do analytical work of any but the simplest sort unless you are a trained chemist.

HOW TO MAKE INDICATORS FOR ACIDS AND ALKALIS

There are a number of substances that chemists call *indicators.* These, in the presence of certain other substances, assume a very deep color, or change from one color to another. Thus, litmus is red in the presence of acids and blue in the presence of alkalis.

SOME COMMON INDICATORS

Indicator	*Color in presence of acids*	*Color in presence of alkalis*
Litmus	Red	Blue
Phenolphthalein	Colorless	Magenta
Methyl orange	Red	Yellow
Congo red	Blue	Red
Cochineal	Orange	Mauve

Litmus, phenolphthalein, and methyl orange are the indicators most frequently used.

Litmus is a dyestuff of unknown chemical make-up which may be purchased in the form of a dry powder or lumps. To make a litmus solution, mix the powder with a little water and boil for some time. When the solution has cooled, filter it. Strips of absorbent paper (mimeograph paper) dipped into this solution while hot will be dyed blue. These strips should be dried and stored in a covered wide-mouthed bottle or a glass jar.

Phenolphthalein is a colorless substance. It should be dissolved in ethyl alcohol and kept in a small bottle fitted with a medicine dropper.

Methyl orange should be dissolved in water.

Congo red should be dissolved in water and used to dye strips of absorbent paper. Paper dipped in Congo red differs from litmus-paper in that it shows graduations in its blue color when dipped in acid, the blue being much more distinct with an active acid such as sulfuric than with a relatively weak one such as boric acid.

Cochineal is sold in liquid form for coloring cake frosting, etc. This should be greatly diluted for chemical purposes.

How to Make Lead Acetate Paper

LEAD ACETATE

Lead acetate paper is used for testing sulfides. The presence of a sulfide turns lead acetate paper black.

Lead acetate paper is made by dipping absorbent (mimeograph) paper in a solution of lead acetate. When the paper has dried, preserve it in a tightly corked bottle where it will be safe from sulfide fumes.

When used to test for hydrogen sulfide, lead acetate paper must be slightly moist.

Litmus Test for Acid

BLUE AND RED LITMUS PAPER VINEGAR

To 3 or 4 teaspoonfuls of water contained in a clean beaker, add 1 teaspoonful of vinegar. Vinegar contains acetic acid. Dip a piece of *red* litmus-paper in the diluted vinegar. The red litmus-paper will not change color. Dip a piece of *blue* litmus-paper in the diluted vinegar. It will turn red. ACIDS turn blue litmus RED.

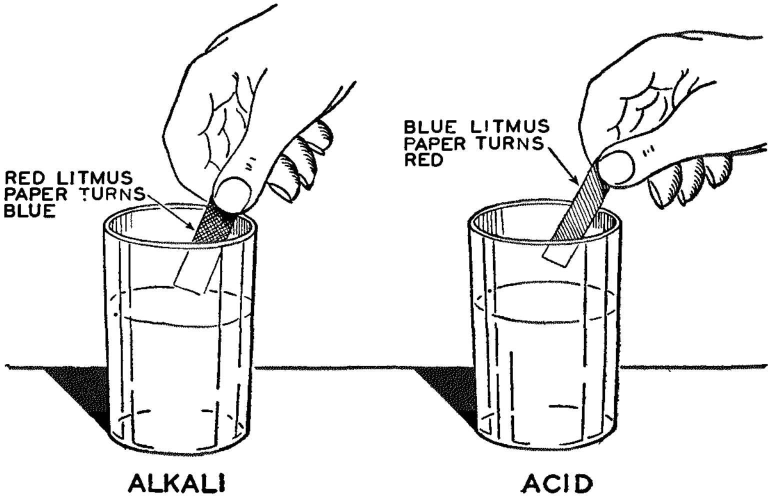
Test solutions of boric acid, lemon juice, tartaric acid, sulfuric acid, etc. Any substance which is an acid or which has an acid reaction when dissolved in water will turn blue litmus-paper red.

Litmus Test for Alkali

BLUE AND RED LITMUS-PAPER SODIUM HYDROXIDE

To 3 or 4 teaspoonfuls of water contained in a clean beaker, add 2 measures of sodium hydroxide. Dip a piece of blue litmus-paper in this solution of sodium hydroxide. It will not change color. Dip a piece of red litmus-paper in the sodium hydroxide solution. It will turn blue. ALKALIS turn red litmus-paper blue.

LITMUS AS AN INDICATOR

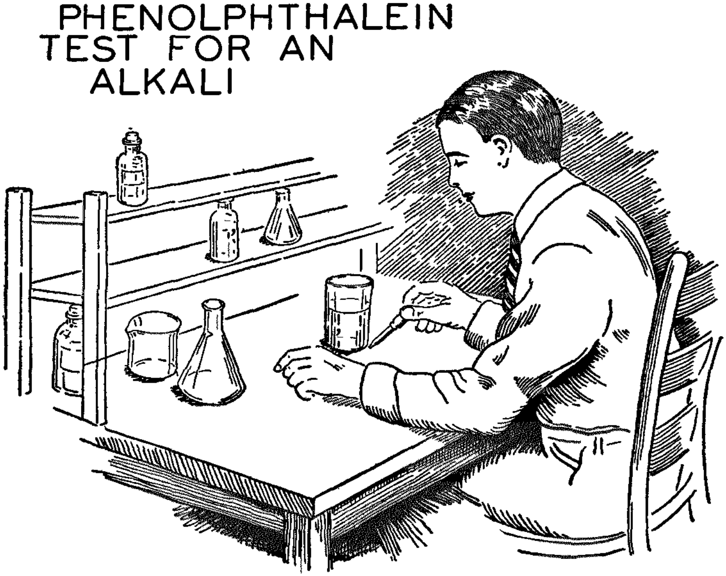


Indicators are substances which, in the presence of certain other substances, assume a deep color, or change sharply from one color to another. Litmus is the most common indicator for acids and alkalis. It is a vegetable dyestuff. Acid turns it red and an alkali or base turns it blue. Litmus paper is unsized paper colored with litmus.

Phenolphthalein Test for Alkali

PHENOLPHTHALEIN VINEGAR SODIUM HYDROXIDE

To 3 or 4 teaspoonfuls of water contained in a test-tube, add 1 teaspoonful of vinegar. Add 2 or 3 drops of phenolphthalein solution. No change will occur. A vinegar solution is a dilute acid. Acid does not change the color of phenolphthalein.



Phenolphthalein is a colorless substance used as an indicator by chemists. In neutral or acid solutions it has no visible color. When an alkali is added to a solution containing phenolphthalein any acid present is first neutralized. Then if there is sufficient alkali to make an excess, the solution will turn red. Phenolphthalein shows best the presence of an excess of an active alkali such as potassium or sodium hydroxide. With a weak base like ammonium hydroxide, a considerable excess of the alkali must often be used before the color appears.

To 3 or 4 teaspoonfuls of water contained in a test-tube, add 2 measures of sodium hydroxide. Shake well to dissolve the hydroxide and add 2 or 3 drops of phenolphthalein solution. It will turn red. ALKALIS turn phenolphthalein MAGENTA (purplish red).

Logwood as an Indicator

LOGWOOD SODIUM BISULFATE SODIUM CARBONATE

Put 2 teaspoonfuls of water in a test-tube and add 6 measures of logwood. Heat this solution gently until it becomes red. Then pour it carefully into another test-tube so as to separate the clear solution from the logwood chips. Save the solution and throw away the chips.

The solution contains a coloring matter called HÆMATOXYLIN, extracted from the logwood. Hæmatoxylin behaves like litmus. It is blue in the presence of alkalis and red in the presence of acids.

Add to the logwood solution 1 measure of sodium carbonate and 2 measures of sodium bisulfate. Notice the change in the color of the logwood solution due to the presence of an acid.

Test for Hydrogen Sulfide

SULFIDE TEST PAPER

Moisten a piece of sulfide test paper and hold it in the mouth of a test-tube in which hydrogen sulfide is being generated. The paper will turn black, due to the formation of LEAD SULFIDE.

Testing a Cabbage for Hydrogen Sulfide

SULFIDE TEST PAPER

Some of the odor of cabbage when it is cooked is due to hydrogen sulfide. Hold a piece of sulfide test paper in the steam issuing from a pot of boiling cabbage. It will turn black due to the formation of LEAD SULFIDE.

Making Cobalt Paper

COBALTOUS CHLORIDE ABSORBENT PAPER

Cobalt is a silver-white metal with a faint suggestion of pink. It has no commercial uses. Cobaltous chloride is a red crystalline compound, which becomes deep blue when partly dehydrated or deprived of its water. A diluted solution of cobaltous chloride makes an interesting invisible or "sympathetic" ink. The writing is almost invisible, but becomes blue when warmed. Afterwards it takes up moisture from the air and becomes invisible again.

Make a solution of cobaltous chloride by dissolving ¼ level teaspoonful in several teaspoonfuls of water contained in a clean beaker.

Cut a few strips of filter paper about ⅜ inch wide and wet each strip by dipping into the cobaltous chloride solution. Pin these up where they will dry. When dried at ordinary room temperature the strips will be PINK to LIGHT BLUE in color.

Dry some of the strips by warming well *above* the flame of a Bunsen burner. Be very careful not to burn the paper. Notice that those strips which are dried over the Bunsen flame are a much DEEPER BLUE than those dried at room temperature. This deeper color is due to more complete drying of the tiny cobaltous chloride crystals with which the filter paper is saturated.

A Test for Water with Cobalt Paper

ETHYL ALCOHOL COBALT PAPER GASOLINE

Pour 2 or 3 teaspoonfuls of water into a dry saucer and then lay a strip of "cobalt" chloride paper in the water. The strip of cobalt paper should be one which has recently been dried over a Bunsen flame and still retains its deep-blue color. If it is exposed to the air for any length of time after being dried, it will sometimes absorb sufficient moisture from the air to affect the color.

When the dry strip of cobalt paper absorbs the water in the saucer it will change in color from deep blue to very pink.

Repeat the test, using a fresh strip of deep-blue cobalt paper and 2 or 3 teaspoonfuls of ethyl alcohol in a *dry* saucer in place of the water. This time the cobalt paper will not change color.

In a third dry saucer pour 2 or 3 teaspoonfuls of gasoline and lay a fresh strip of dark-blue cobalt paper in it. The paper will soak up the liquid but there will be no change in its color.

It is WATER only that causes the dark-blue cobalt paper to change to pink. Therefore cobalt paper and its change from blue to pink may serve as a TEST FOR WATER.

Tests for Water in Animal, Vegetable, and Mineral Matter

COBALT PAPER APPLES POTATOES

Put some small pieces of raw apple in a clean test-tube and push them to the bottom of the tube with a glass stirring rod. Use a test-tube clamp to hold the tube. Hold the lower end of the tube containing the pieces of apple in the tip of the flame of an alcohol lamp or Bunsen burner. Keep the mouth of the tube lower than the end held in the flame. The precaution will prevent any liquid that is formed and condenses on the walls of the test-tube from running back into the hot end of the tube and cracking it. When you notice that a liquid is forming on the inner walls of the cooler end of the tube, test this liquid by thrusting a strip of dark-blue cobalt paper into the tube. The change in the color of the paper from dark blue to pink proves that the liquid is WATER.

Test some pieces of raw potato, carrot, or other vegetables and fruits in the same way. You will find that they all contain water which is driven forth when they are heated. Tomatoes, asparagus, lettuce, celery, strawberries, and cabbage are more than 90 per cent WATER by weight.

Test a piece of lean beef in place of the fruits and vegetables. When heated, a vapor will be driven forth and condense on the inner walls of the cooler end of the test-tube. A piece of dark-blue cobalt paper will prove that this liquid is WATER.

If you can obtain some gypsum (rock), heat a few small pieces about as large as peas in the test-tube and test the liquid which is driven forth with a piece of dark-blue cobalt paper. The change in color from blue to pink will show that gypsum contains water.

The chemical name for gypsum is hydrated calcium sulfate. When the water is driven out by heating, it is ANHYDROUS or DEHYDRATED. Anhydrous calcium sulfate is known commercially as PLASTER OF PARIS. The white powder remaining in the test-tube in the last experiment is plaster of Paris.

Testing Water for Lime

SODIUM CARBONATE

Put the water to be tested in a test-tube and add a small amount (about 2 measures) of sodium carbonate. Shake up the solution and allow it to stand for 15 minutes. If the water contains lime it will become cloudy due to the formation of calcium carbonate or chalk.

Testing Water for Iron

SODIUM FERROCYANIDE

Well water and spring water sometimes contain an appreciable amount of iron in solution. Half fill a test-tube with some of the water to be tested and add a few small crystals of sodium ferrocyanide. Shake until the crystals have dissolved and then let the tube stand in the rack for a few minutes. If a blue precipitate forms, there is iron in the water.

Green Test for Boron Compounds

BORIC ACID ALCOHOL

Boron is an element. It is found chiefly in the form of its acid, known as boric acid, and in the form of salts of boric acid. Borax is one of the salts of boric acid. Under the name of boracic acid, boric acid is used in medicine as an antiseptic solution which is especially valuable as an eye-wash. When boric acid is burned with alcohol, it gives a beautiful green color to the resulting flame. Since an alcohol flame is ordinarily colorless, the presence of green in the flame may be used as a test for boron compounds.

Dissolve a small pinch of boric acid in a test-tube one-half full of alcohol. Pour a few drops of the solution in a spoon and light it. It will burn with the green flame which is characteristic of boron compounds.

A Chemical Test for Copper

COPPER FILINGS NITRIC ACID HOUSEHOLD AMMONIA

File a copper coin or wire until you have a measure full of filings. Place these in a clean, dry test-tube. Then, using a medicine dropper, add 10 drops of concentrated nitric acid.

When concentrated nitric acid comes in contact with copper, CUPRIC NITRATE and almost pure NITROGEN TETROXIDE are formed. Copper and somewhat diluted nitric oxide produce cupric nitrate and NITRIC OXIDE. Nitrogen tetroxide and nitric oxide are both gases. That they are both produced by the reaction of copper and nitric acid is a good example of the fact that the interactions of metals with nitric acid are usually complex. Platinum and gold alone of all the metals are not attacked by nitric acid.

After the nitric acid has had an opportunity to work on the copper for 10 minutes or so, add 10 drops of water to the contents of the test-tube. Then add household ammonia, a few drops at a time, until, after shaking, the mixture has a decided odor of ammonia.

The liquid in the test-tube will now have a decidedly BLUE color. This is one of the chemical tests by which copper is recognized. Solutions of nickel salts also produce a blue color when mixed with household ammonia

A Test for Copper

POTASSIUM FERROCYANIDE   CUPRIC SULFATE

Put 1 measure of ferrocyanide of potassium into a test-tube containing 2 or 3 teaspoonfuls of water. Pour some of this solution into another test-tube containing a dilute solution of cupric sulfate. A gelatinous brown precipitate of CUPRIC FERROCYANIDE will form.

Potassium ferrocyanide may be used as a test to detect the presence of copper.

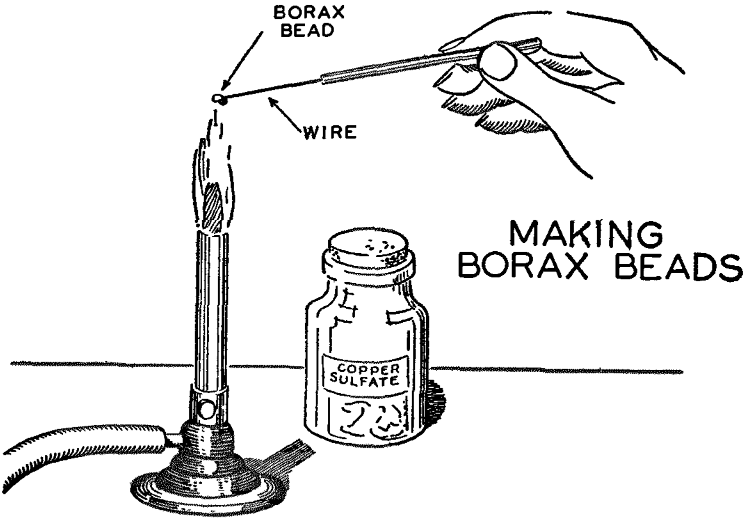
Test for Nickel

If you want to find out if the plating on a small metal object is nickel, put the article to be tested into a small beaker and cover it with water. Add a teaspoonful of sodium bisulfate to the water and heat it for a few minutes. If the solution becomes green, the plating is nickel.

Borax Bead Tests

SALTS OF COPPER COBALT NICKEL IRON AND MANGANESE

When heated, borax loses its water of crystallization and swells up into a white porous mass. Then it melts to a clear glass—borax glass—which is anhydrous borax (borax without water).



When heated with oxides of metals, borax (sodium tetraborate) forms a colored mass known as a bead. The bead is colored in various tints according to the metal used. Thus with a trace of copper, the bead is colored green. Cobalt gives a deep-blue color.

Fused borax dissolves metallic oxides, and these often impart to the glassy mass a color characteristic of the metal. Thus, when heated with a drop of fused borax, manganese compounds give a violet color, cobalt compounds a blue color, etc., as shown in the following table.

COLORS OF BORAX BEADS PRODUCED IN OXIDIZING FLAME OF BUNSEN BURNER

Metal	*Color*
Copper	Green if hot; blue if cold
Cobalt	Blue if hot or cold
Chromium	Green or red if hot or cold
Iron	Yellow if cold; brown if hot
Nickel	Violet if hot; brown if cold
Manganese	Amethyst if hot or cold

In order to make borax bead tests, a Bunsen burner and a piece of platinum wire are necessary.

Cut a piece of glass tubing about 4 inches long. Heat one end of this tube until it is nearly closed and seal a piece of fine platinum wire about 2 inches long in the red-hot tube. The glass tube attached to the wire will serve as a handle. When it has cooled, twist the free end of the wire into a loop around the wooden part of a match.

Heat the wire until red in the upper part of a Bunsen burner flame and dip the red-hot loop into powdered borax. Some of the borax will adhere to the loop. Heat again until the borax is melted into a small, transparent, colorless bead. If the bead is not large enough to fill the loop, dip it into borax a second time and melt again.

Heat a colorless, transparent, borax bead to redness and then touch it to a small piece of copper sulfate crystal about the size of a pin head. The hot bead will pick up the crystal. Fuse the two together by heating in the upper portion of a Bunsen burner flame for several minutes. The bead will become green. When it is removed from the flame and has cooled, the bead will be blue.

Open the wire loop and detach the bead. Clean the wire and form it into a loop again by bending it around a wooden match. Make a new bead (a new colorless and transparent bead is necessary for each test) and test the effect of a small crystal of cobaltous chloride about as large as a pin head when fused with the bead. You will find that it will color the bead dark-blue.

Make other tests, using salts of chromium, iron, nickel, and manganese.

A Test for Ozone

POTASSIUM IODIDE

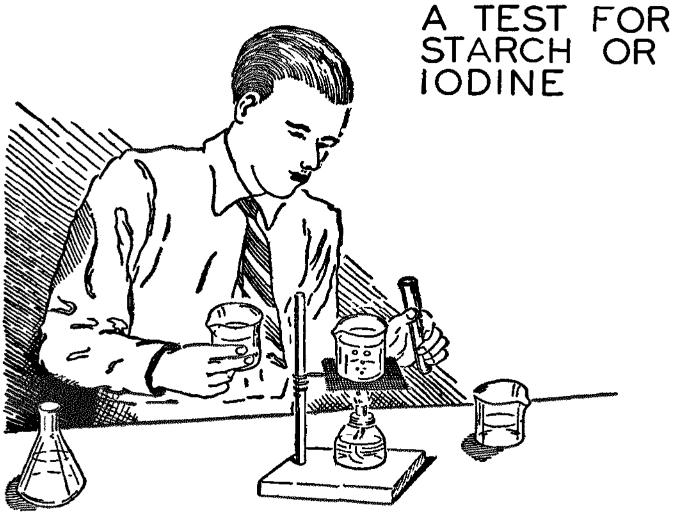
Unlike oxygen, ozone will react with potassium iodide, liberating IODINE, WATER, and POTASSIUM HYDROXIDE. This fact can be used as a test for ozone. Dip a strip of absorbent paper into a solution of potassium iodide and hold it in the mouth of the beaker in which ozone is being generated (see page 101.) The paper turns brown, owing to the liberation of iodine.

If a little starch is mixed in the potassium iodide solution, a piece of moist paper which has been dipped in it will be colored blue when exposed to ionized air.

A Test for Starch or Iodine

Prepare some starch paste by boiling 5 tablespoonfuls of water in a beaker glass and stirring into it ½ teaspoonful of laundry starch previously reduced to the consistency of cream by thoroughly mixing it with a spoonful or so of water.

Put a teaspoonful of this starch paste into a clean test-tube and add water until the test-tube is half filled. Add 1 or 2 drops of tincture of iodine. A deep-blue color will be produced. If heated, the blue color will disappear and reappear upon cooling.



The behavior of free iodine and starch toward each other furnishes an accurate test for both substances. They produce a deep blue color.

The blue substance in the test-tube is not a chemical compound but a solution of the iodine in the little starch particles suspended in the water.

The behavior of free iodine and starch toward each other furnishes an accurate test for both substances.

A Test for Sugar

SUGAR	SODIUM CARBONATE
CALCIUM OXIDE	COBALTOUS CHLORIDE

When we say "sugar" we usually mean the familiar white granulated sugar used on the dining table to sweeten tea, coffee, etc. To the chemist this is SUCROSE and it is only one of a number of sugars, the best known of which are SUCROSE, FRUCTOSE, and GLUCOSE or DEXTROSE.

Most of our sucrose is obtained from sugar-cane. It also occurs in certain fruits and vegetables and in maple sap.

Fructose is found in ripe fruits and in honey.

Glucose, also known as grape sugar and dextrose, is most familiar to us in the form of "Corn" sirup. It is made commercially by boiling starch obtained from corn in water containing a small amount of hydrochloric acid.

Put 3 measures of granulated sugar in a test-tube about ⅓ full of water. Shake the tube thoroughly so as to dissolve the sugar.

In a second test-tube about ⅓ full of water, put 1 measure of calcium oxide (quicklime), 1 measure of sodium carbonate, and ½ measure of cobaltous chloride. Shake until dissolved, then pour this solution into the sugar solution.

If the sugar is pure granulated sucrose, the solution will be an amethyst-violet color.

Add 1 or 2 drops of corn sirup, shake thoroughly, and the color will change to light green or turquoise-blue.

A Test for Dextrose

COPPER SULFATE

Make a dilute solution of copper sulfate and pour enough in a test-tube to half-fill it. To this solution add a few drops of water in which some cane sugar or sucrose has been dissolved. Heat the mixture. No change will be noticed.

To another test-tube half-filled with a dilute solution of copper sulfate add a few drops of corn sirup. Heat the mixture and it will change in color from blue to yellowish-green. This is a rough test for dextrose or glucose.

For an accurate test for dextrose, Fehling's or Haine's solution is used. Fehling's solution is an interesting reagent containing copper, used for the analysis of sugars and the detection of the organic chemicals known as *aldehydes.* When Fehling's solution is added to a solution containing dextrose, a precipitate of copper oxide is formed.

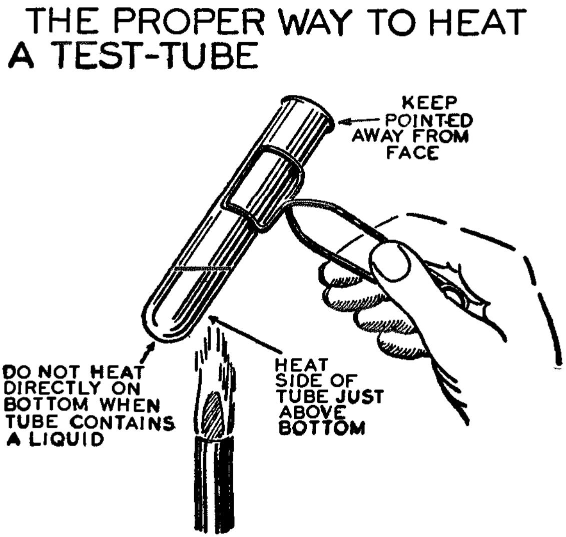
Fehling's Solution, a Test for Dextrose

CREAM OF TARTAR BAKING POWDER
SODIUM CARBONATE
CORN SIRUP
COPPER SULFATE
CALCIUM OXIDE
MAPLE SIRUP

Maple sirup is frequently adulterated with cane sugar and dextrose. The cane sugar and dextrose are not harmful. They are added because they are cheaper than pure maple sirup. You can discover the presence of dextrose in maple sirup with the aid of Fehling's solution.

To a test-tube half-filled with water, add a little corn sirup. Add Fehling's solution and heat. The precipitate which forms is COPPER OXIDE. The dextrose in the corn sirup caused the precipitate.

To a second test-tube half-filled with water, add a little of the maple sirup which is to be tested. Add some Fehling's solution and heat. If a precipitate of copper oxide forms, the maple sirup contains dextrose.



Safety first should be the motto of every young chemist. A long step in the right direction is the formation of correct habits in handling chemicals and chemical apparatus. When heating a test-tube, keep the open end pointed away from your face. If the test-tube contains a liquid, do not heat the bottom of the tube. Heat the side, just above the bottom. This precaution will prevent liquid contents from being blown suddenly out of a tube.

Fehling's solution can be purchased at most drug-stores. It is used by physicians for making certain pathological tests. Fehling's solution is made by mixing a solution of copper sulfate with one containing sodium hydroxide and Rochelle salt. Rochelle salt is potassium and sodium tartrate. It is used in baking powder. If the ingredients are in proper proportion, copper is not precipitated as copper hydroxide as it is ordinarily when solutions of sodium hydroxide and copper sulfate alone are mixed. Instead, because of the presence of Rochelle salt, the copper forms a deep-blue compound with the Rochelle salt.

You can make your own Fehling's solution for the corn sirup and maple sugar tests.

To a test-tube half-full of water, add 3 measures of cream of tartar baking powder. When the reaction stops, add 1 measure of copper sulfate and shake until dissolved. Then add 4 measures of sodium carbonate and 4 measures of calcium oxide (unslaked lime). Shake thoroughly and filter the mixture. The liquid which passes through the filter is Fehling's solution.

A Test for Dichromic Acid

POTASSIUM DICHROMATE HYDROGEN PEROXIDE
SULFURIC ACID

The presence of a dichromate in a solution may be detected by a test using hydrogen peroxide and sulfuric acid.

Put 4 or 5 teaspoonfuls of hydrogen peroxide in a small beaker. Add enough sulfuric acid, drop by drop, to make the hydrogen peroxide slightly acid. As the acid is added slowly, test from time to time with blue litmus-paper, and as soon as the blue paper turns red, stop adding acid.

If a few drops of a solution containing a dichromate are added to the acidulated hydrogen peroxide, a beautiful blue color will be formed. You can prove the test by adding a drop of a solution of potassium dichromate.

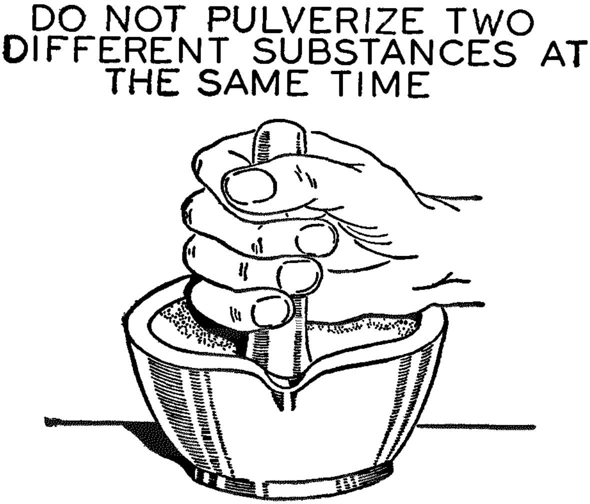
CHAPTER VII

SAFE "FIREWORKS"

THE bombs, rockets, and elaborate displays used to celebrate the Fourth of July are made of chemicals. They are made by trained men who know how to take every precaution to guard against explosions. Even so, the manufacture of commercial fireworks is a dangerous business. Don't ever be so foolish as to try to make real fireworks. It is not worth the risk involved.

The "fireworks" described in this chapter are not real fireworks. They are experiments in combustion which are perfectly safe if you follow directions explicitly and literally. Do not use larger quantities than the experiments direct. When the instructions say *mix,* do exactly that and do NOT *grind* or *rub.* When the instructions say grind *separately,* do NOT attempt to grind two substances together to see what happens.

Some of these experiments should be performed OUTDOORS. They have each been labeled "An OUTDOOR Experiment." They do not produce much fire or flame on account of the small quantities of chemicals used, but in some instances do make a cloud of smoke having the familiar "fireworks odor."



The friction between two substances rubbed together in a mortar sometimes will bring about a violent chemical reaction. Avoid the risk of an explosion by using only a clean mortar and pestle and pulverizing only one substance at a time.

Green Flame

BORIC ACID   ALCOHOL

Ordinarily, alcohol produces a colorless flame.

Dissolve some boric acid in alcohol and pour it into a clean porcelain crucible or evaporating dish. Ignite the solution and stir it with a glass rod. The flame will be a beautiful green color.

Making Touch Paper

POTASSIUM NITRATE OR SALTPETER

Dissolve 3 teaspoonfuls of potassium nitrate in ¼ cup of water. This will make a strong solution. Cut some mimeograph paper into strips about ⅜ to ½ inch wide and soak them in this potassium nitrate solution. Then hang them up to dry.

When dry, ignite the end of one of the strips. It will not blaze but will sizzle and smolder until completely consumed. The oxygen consumed in the combustion of the paper is supplied by the potassium nitrate.

How to Make a Fuse

POTASSIUM NITRATE SOLUTION

Soak a piece of white cotton string in the potassium nitrate solution prepared in the previous experiment. Then hang it up to dry. When dry, ignite one end of the string. It will sputter and sizzle and burn like the fuse of a fire-cracker.

Sham Battles

POTASSIUM NITRATE SOLUTION

Draw the outlines of bombers, battle-ships, submarines, forts, etc., on a sheet of mimeograph paper. Using an artist's camel's hair brush and a strong solution of potassium nitrate as ink, draw a line between the battle-ship and fort, another line between the plane and "sub," etc., to represent the paths of the projectiles which they are firing at each other.

When dry, the lines will be invisible. But if ignited with a glowing spark, they will sputter and glow, showing the path of make-believe projectiles much as tracer bullets do in real warfare.

Magnesium Flare

MAGNESIUM RIBBON

Hold a piece of magnesium ribbon with a pair of pliers or tongs, and light it in the flame of a Bunsen burner or alcohol lamp. It will take fire more readily there than it will from a match.

The ribbon will burn and give off a light of dazzling brightness. The white smoke and the grayish powder that remain are MAGNESIUM OXIDE.

The parachute flares which are dropped from airplanes at night, when it is necessary to make an emergency landing, burn magnesium.

White Flashlight Powder

An OUTDOOR Experiment

POWDERED MAGNESIUM   POTASSIUM NITRATE

The light of burning magnesium is rich in waves of short lengths which are useful in photography. The bright light of photographer's flashlights is produced by burning magnesium.

Magnesium is also used in fireworks. During the World War (1914-1918) magnesium was used in large quantities in making tracer bullets, star shells, and flares. It is also one of the ingredients of incendiary bombs used in warfare.

Mix 2 measures of powdered magnesium with 2 measures of potassium nitrate. Do not rub or grind these substances together.

Put the mixed powders in a heap on a stone or in an old crucible. Embed in the powder one end of a fuse made by soaking a soft cotton string in potassium nitrate solution. The fuse should be dry before it is used.

Light the fuse and stand away. When the burning fuse reaches the powder, there will be a brilliant flash. The potassium nitrate furnishes oxygen to the magnesium so that it burns very quickly.

Aluminum Flash

An OUTDOOR Experiment

ALUMINUM POWDER POTASSIUM DICHROMATE

Mix ½ teaspoonful of powdered aluminum (the aluminum bronze used in paint) with ½ teaspoonful of powdered potassium dichromate by rubbing the two substances gently together in a mortar.

Put the mixture in a small heap on a flat stone or brick *outdoors*. Place the heads of three or four matches on top of the pile and ignite them by touching them with the red-hot end of an iron rod. Or you can ignite the mixture by embedding one end of a fuse made of a soft cotton string which has been soaked in potassium nitrate solution and dried.

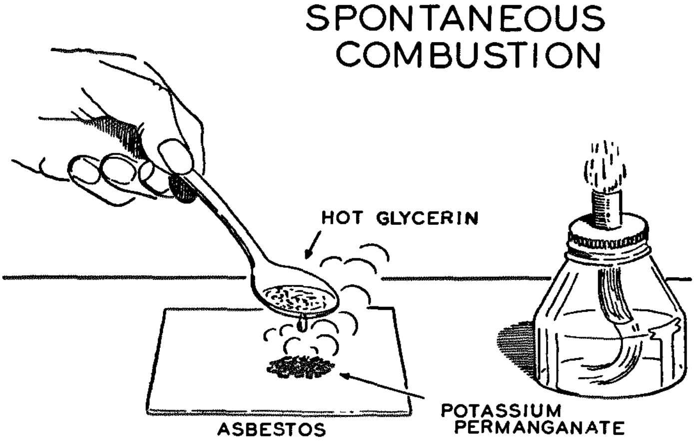
The mixture of aluminum and potassium dichromate will burn with a brilliant flash. Do not lean over the mixture when you fire it.

This experiment is not dangerous if you do not use more than 1 teaspoonful of the mixture at a time.

Spontaneous Combustion

POTASSIUM PERMANGANATE GLYCERIN

Grind ½ teaspoonful of potassium permanganate to a fine powder in a clean mortar. Make a pile of the powder on a flat stone or a thick piece of asbestos.



Spontaneous combustion is an active burning started by the accumulation of heat produced by oxidation. In the experiment illustrated above, the oxidizing process is speeded by using warm or heated glycerin. The warm glycerin is oxidized so rapidly by the potassium permanganate that it takes fire.

Warm a few drops of glycerin in a spoon held *over* the flame of an alcohol lamp or Bunsen burner. Pour the hot glycerin on the powdered permanganate. The glycerin will take oxygen away from the permanganate so rapidly that it will become very hot and catch on fire.

Coloring the Flame of a Bunsen Burner

CALCIUM CHLORIDE	BARIUM CHLORIDE
STRONTIUM NITRATE	POTASSIUM NITRATE
CHARCOAL	IRON WIRE

For this experiment you will need four pieces of ungalvanized iron wire about 10 inches long and four pieces of charcoal the size of a small marble.

Twist one end of each wire around a piece of charcoal. Using the wire as a support, hold one of the pieces of charcoal in the flame of a Bunsen burner. Observe that the burning charcoal gives the flame a slightly yellowish tinge.

Remove this piece of charcoal from the flame and moisten it in a solution of calcium chloride. Put it back in the flame. This time it will impart a REDDISH-YELLOW tinge.

Dip the second piece of charcoal in a solution of barium chloride and hold it in the flame. It will color the flame GREEN.

Moisten the third piece of charcoal in a solution of strontium nitrate and hold it in the Bunsen flame. The color will be a beautiful CRIMSON.

Moisten the fourth piece of charcoal with a solution of potassium nitrate. Its effect upon the Bunsen flame will be to give it a VIOLET color.

Various chemical salts are used in the manufacture of fireworks and pyrotechnics to produce colors.

Red Flash Powder

An OUTDOOR Experiment

POWDERED MAGNESIUM   POWDERED STRONTIUM NITRATE

On a sheet of paper, mix, without grinding or rubbing, 2 measures of powdered magnesium and 2 measures of strontium nitrate.

Pour the mixture into a little pile on a flat stone or asbestos shingle and ignite it by using a fuse, as in the previous experiment.

There will be a sudden, brilliant flash of RED light. The burning strontium furnishes the color.

Green Flash Powder

An OUTDOOR Experiment

POWDERED POTASSIUM NITRATE   BORIC ACID
POWDERED SULFUR   MAGNESIUM POWDER

On a sheet of paper, mix carefully, but *do not rub or grind* them together, 2 measures of powdered potassium nitrate, 2 measures of powdered boric acid, 2 measures of powdered magnesium, and 2 measures of powdered sulfur.

Pour this mixture in a little pile on a flat stone or asbestos shingle and ignite it by using a fuse. It will burn with a brilliant GREEN flash.

Red Fire

An OUTDOOR Experiment

POWDERED CHARCOAL
POWDERED SULFUR
POWDERED POTASSIUM NITRATE
POWDERED STRONTIUM NITRATE

On a sheet of paper, mix, but *do not grind,* 4 measures of powdered charcoal, 2 measures of powdered sulfur, 4 measures of powdered potassium nitrate, and 2 measures of powdered strontium nitrate.

Pour this mixture into a small pile on a flat stone or an asbestos shingle and ignite it with a match or a piece of fuse. The mass will take fire and burn with a RED light. The red color is due to strontium. This mixture does not burn as rapidly as the flash powders.

Yellow Fire

An OUTDOOR Experiment

POWDERED CHARCOAL
SODIUM CHLORIDE (SALT)
POWDERED SULFUR
POWDERED POTASSIUM NITRATE

On a sheet of paper, mix, but *do not grind,* 4 measures of powdered charcoal, 2 measures of sodium chloride, 4 measures of powdered sulfur, and 4 measures of powdered potassium nitrate.

Pour this mixture into a small pile on a flat stone or asbestos shingle and ignite it with a match or a piece of fuse. The powder will take fire and burn with a YELLOW light. The yellow color is due to sodium.

Green Fire

An OUTDOOR Experiment

POWDERED CHARCOAL
POWDERED SULFUR
POWDERED POTASSIUM NITRATE
POWDERED ZINC

On a sheet of paper, mix, but *do not grind,* 4 measures of powdered charcoal, 2 measures of powdered sulfur, 4 measures of potassium nitrate, and 4 measures of powdered zinc.

Pour this mixture into a small pile on a flat stone or asbestos shingle and ignite it with a match or piece of fuse. The powder will take fire and, due to the presence of zinc, will burn with a GREEN flame.

Rainbow Fire

An OUTDOOR Experiment

Prepare mixtures for making green, red, and yellow fire as described in the last three experiments. Keep them separate from one another.

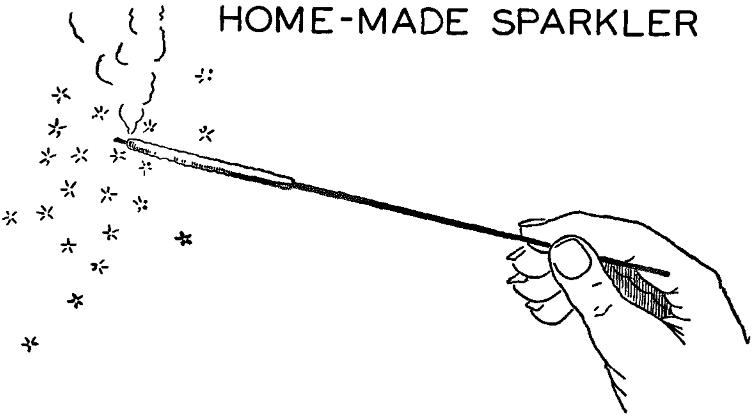
When ready to burn them, pour each mixture in a pile on an asbestos shingle or flat stone. The piles should be in a row and just touch one another. Put one end of a fuse in one of the mixtures and ignite it. The different mixtures will burn and produce light of three different colors.

Sparkling Matches

PARAFFIN
POWDERED MAGNESIUM
POWDERED POTASSIUM NITRATE

On a sheet of paper, mix, but *do not grind,* 4 measures of powdered magnesium and 2 measures of potassium.

Melt some paraffin in a spoon or an evaporating dish and dip the heads of some matches in it so as to coat the match sticks for about ½ inch. Roll each match, while the paraffin on it is still hot, in the mixture of powdered magnesium and potassium nitrate so that it picks up a coating of the powders. Allow the matches to dry and harden. When the matches are lighted they will give out bright flashes and sparks.



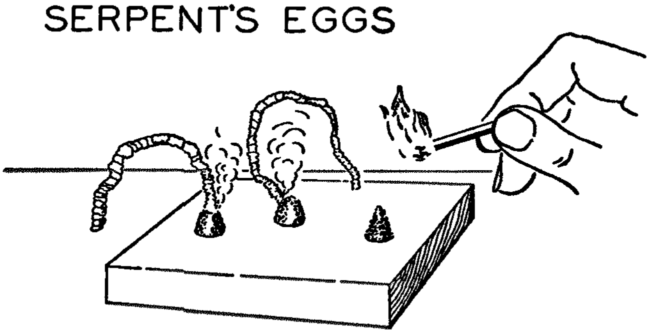
If, instead of matches (see experiment called "Sparkling Matches") long, thin wooden splints are used, they may be coated with the combustible mixture of paraffin, potassium nitrate and magnesium for more than ½ inch. They will burn and sparkle much longer than the matches.

Serpent's "Eggs"

POTASSIUM NITRATE SUGAR POTASSIUM DICHROMATE

When lighted, these "eggs" will burn and a brown snakelike ash with a length of several feet will wriggle forth.

Mix carefully, *without* rubbing or grinding, 1 teaspoonful of potassium nitrate with 1 teaspoonful of powdered sugar and 2 teaspoonfuls of powdered potassium dichromate. Then add, drop by drop, just enough mucilage to the mixture to make a dough which can be molded with your fingers like modeling clay.



When "serpent's eggs" burn a long wriggling ash called "Pharaoh's Serpent" sprouts forth.

Press the mixture into little cone-shaped pieces the size of large peas and let them dry thoroughly. When the cones are dry and lighted at the tip, the "serpents" will writhe out.

CHAPTER VIII

EXPERIMENTS WITH A FEW ORGANIC COMPOUNDS

ESTIMATES place the number of different chemical compounds at the amazing total of 250,000 to 500,000. Approximately 225,000 are known to chemists. They are divided into two groups.

The larger group, numbering more than 200,000 compounds called ORGANIC compounds, includes all but a few of those substances which contain CARBON.

The compounds of all the other elements except carbon are called INORGANIC compounds. They total about 25,000 in number. A few carbon products, carbon dioxide and carbon monoxide among them, are included in the inorganic substances.

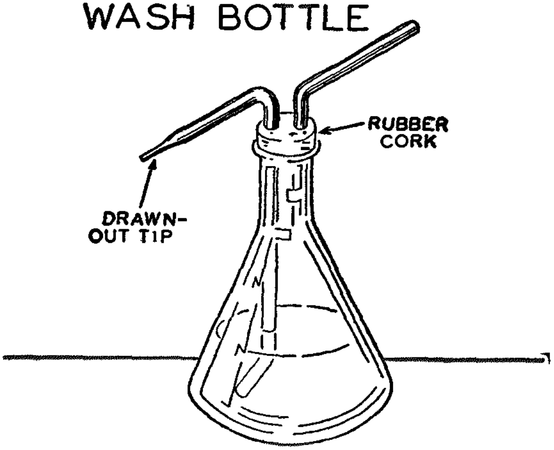
Most of the chemist's creative work, the production of new materials, medicines, paints, plastics, fabrics, and whatnot which make our modern industry and present-day standards of living possible, belong to the realm of organic chemistry.

Extracting Casein from Milk

MILK VINEGAR

Milk is often called the "perfect food," particularly for children. Aside from being 87 per cent water, it has practically all the essential food elements in the proper proportion for health.

In addition to water, butter fat, mineral salts, and milk sugar, cow's milk contains about 3 per cent of the protein matter called CASEIN. Casein is a white to yellowish solid suspended in the water of the milk in a finely divided state. When milk sours, the casein changes into a spongy mass called CURD, which floats on the water. The liquid part of the milk is then called WHEY.



A wash bottle is handy when small quantities of water are required in the laboratory. Blowing air into the bottle through the upward pointing tube forces a thin stream of water out of the jet.

Often in a chemical laboratory it is necessary to wash a precipitate after it has been filtered so as to free it from the filtrate. A jet of water from a wash bottle directed upon the precipitate as it rests upon the filter paper in the funnel will do the job. You can make a wash bottle out of an Erlenmeyer flask, a two-hole cork and glass tubing.

Milk is not only a food and a source of butter and cheese; it is a raw material of industry. Skim milk is treated with acid to precipitate the casein, and the casein is employed in making molding materials, glue, glazing for paper, and imitation amber and ivory. Buttons and umbrella handles are frequently made from casein.

Put about 1 pint of sweet milk into an agate-ware saucepan and heat it very gently to a temperature of 90 degrees Fahrenheit. The temperature is so important that it is necessary to use a thermometer and not to overheat. A 10-cent thermometer from the dime store will be satisfactory.

When a temperature of 90 degrees Fahrenheit has been reached, remove the saucepan from the fire and while stirring continuously with a glass rod pour ½ cupful of vinegar into the milk.

The milk will curdle immediately and, after continued stirring, the curd or CASEIN will gather into a mass which can be lifted out and placed in a dish. By squeezing this rubbery mass like a sponge, you can press most of the water out of it.

The casein manufactured by the butter and cheese industries as a by-product is dried and reaches the market as a granular powder or in small lumps. Argentina is the greatest producer of casein.

Testing Casein for Protein

Proteins are extremely complex chemical compounds which contain nitrogen, carbon, hydrogen, and oxygen. Proteins are found in meat, milk, eggs, fish, beans, peas, and wheat. The presence of protein in a food is easily discovered by treating with concentrated nitric acid. If protein is present a yellow color results. Then, if the food is washed to remove the acid and treated with ammonium hydroxide, an orange color is produced.

Put about ½ teaspoonful of the casein separated from milk in the previous experiment into a large test-tube. Add 3 to 4 cubic centimeters of concentrated nitric acid and heat gently until the acid boils. Continue to boil gently for 2 or 3 minutes. Keep the mouth of the test-tube turned away from your face while boiling acid.

Pour off the acid when the boiling is completed and retain the casein. Its color has been changed to YELLOW by the action of the hot nitric acid. This change in color by nitric acid is a characteristic of protein matter and proves that casein contains protein. It is, in fact, protein.

Further proof that casein is protein may be secured by pouring some ammonium hydroxide into the test-tube after the hot nitric acid has been poured out. The casein will change in color from yellow to orange.

Extracting Starch from a Potato

Rub a clean potato on a grater until you have produced about ½ cupful of pulp. Mix this pulp with water and squeeze through a piece of cotton cloth. Repeat this operation several times. The water or FILTRATE squeezed out of the potato pulp contains starch. The woody fiber or cellulose of the potato remains in the cloth.

Put a small amount of the filtrate in a test-tube and add 2 or 3 drops of tincture of iodine. The blue color which is produced indicates the presence of starch.

Allow the rest of the filtrate to stand until the starch which is suspended in it has settled. Then drain off the water and dry the residue. The result will be starch, a white powder.

Starch is a CARBOHYDRATE found in wheat, corn, potatoes, in the roots and stems and fruits of many plants. It is formed from carbon dioxide and water in the leaves of plants by that important process called PHOTOSYNTHESIS.

Starch is one of our most important foods. It is used in the laundry for stiffening clothes. It is also an essential raw material in the manufacture of glucose, alcohol, and dextrin. Glucose is an important sugar. Dextrin is used in making adhesives. Alcohol is one of the important solvents used in medicine and industry.

Making Dextrin

LAUNDRY STARCH

Heat a teaspoonful of laundry starch in a porcelain dish. Stir constantly. It will turn brown gradually. Continue heating for five to ten minutes. Then add 4 or 5 teaspoonfuls of water and boil gently. A solution of DEXTRIN will be obtained. This can be separated from the starch with which it is mixed by filtering.

Pour some of the solution of dextrin obtained in the previous experiment into a test-tube and add twice its bulk of alcohol. The dextrin will be precipitated. It is insoluble in alcohol.

Dextrin is a white powder, tasteless and odorless. When mixed with water, it forms a sirupy liquid which has strong adhesive qualities. It is used in adhesive pastes, particularly for postage-stamps and envelopes.

Extracting Gluten from Flour

WHITE WHEAT FLOUR

Wet ½ cupful of white wheat flour with enough water to make a thick dough. Wrap the dough in a linen or cotton cloth and knead it under a slow stream of cool water from the faucet until the water runs clear. Small particles of starch squeezed through the cloth will make the water turbid at first. If some of the turbid water is allowed to stand until the starch settles, you can prove that it is starch by means of a few drops of tincture of iodine. The iodine will form a blue stain where it comes into contact with the starch.

The tough, viscous mass remaining in the cloth is GLUTEN, one of the principal constituents of wheat flour.

How to Make Vegetable Parchment

SULFURIC ACID FILTER PAPER

Wood, vegetable fibers, and plant stems are mainly cellulose. Cellulose is a carbohydrate. Linen and cotton are nearly pure cellulose. A great variety of very valuable materials is produced by the reaction of cellulose with certain chemicals. Collodion, guncotton, pyroxylin, celluloid, cellulose, acetate, artificial silk, and cellophane are some of them. Paper is an important cellulose product which is manufactured from paper or rags. The finest kind of filtering paper is nearly pure cellulose.

Strong alkalis decompose cellulose. The various acids have different effects upon it.

Cellulose is converted into a semi-transparent tough substance, resembling animal membrane, by a short contact with sulfuric acid of the right strength.

Tie a string or a rubber band around the middle of a test-tube. Fill the tube with water up to this mark and pour it into a soup plate. Then fill the tube up to the same mark with strong sulfuric acid and pour it *slowly* into the water in the soup plate. Add the same amount of sulfuric acid two more times. You will now have three times as much sulfuric acid in the plate as water. Be certain to pour the *acid into the water* and not the water into the acid or you will be in trouble because the acid will spatter.

When the mixture has become perfectly cold, lay a piece of filter paper in the diluted acid. Let it remain there for 10 seconds. Then remove it and immediately lay it in a pan full of water. Move it around so as to rinse it and remove the acid. Then wash it in some water containing a little household ammonia. Finally, wash it with pure water.

If the paper has not been changed into a parchment-like material, try another piece of filter paper. Immerse it in the acid solution for 15 seconds and then wash it in the manner already described. Keep on experimenting with fresh pieces of filter paper, varying the *time* of immersion slightly until a good result is obtained.

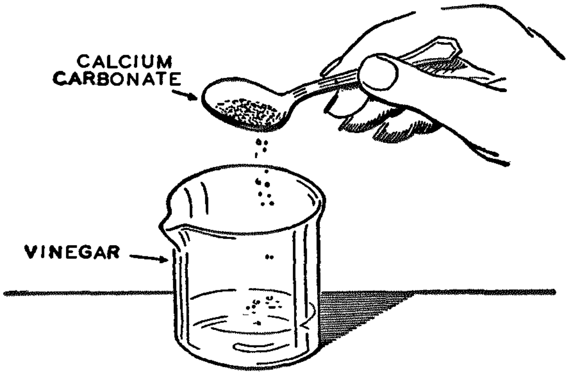
An Experiment with Vinegar

VINEGAR CALCIUM CARBONATE

There are a great number of *organic* acids. Among those which are important for industrial purposes are acetic, tartaric, citric, lactic, butyric, oleic, palmitic, and stearic acids.

Acetic acid is the most important organic acid. It is employed in the production of vinegar, white lead, artificial silk, and photographic film. Vinegar commonly contains from 5 to 13 per cent of acetic acid.

AN EXPERIMENT WITH VINEGAR



This experiment illustrates the reaction between two common substances, vinegar and chalk. One is an acid and the other a carbonate. Vinegar is principally dilute acetic acid. Chalk is calcium carbonate. When calcium carbonate is added to vinegar, carbon dioxide and calcium acetate are formed. The carbon dioxide escapes in the form of bubbles and the calcium acetate remains in solution.

Put 6 tablespoonfuls of strong vinegar into a small beaker glass. Add calcium carbonate (powdered chalk), a little at a time, until there is no more effervescence (escape of bubbles). If you do this slowly and carefully, CALCIUM ACETATE will be formed and remain dissolved in the liquid.

Filter this solution and then evaporate it slowly in an evaporating dish or saucer. The solid residue which is left is impure calcium acetate.

How to Make Acetic Acid

CALCIUM ACETATE SULFURIC ACID

Put some of this calcium acetate into a small test-tube and add a few drops of undiluted sulfuric acid. Heat gently. Calcium sulfate and acetic acid will be formed. Acetic acid may be recognized by its peculiar odor.

Palmitin and Olein

OLIVE OIL

Oils and fats occur in animal tissues and in some plants. Butter, lard, tallow, and cod-liver oil are valuable animal fats. Among the more common vegetable oils used as food are those obtained from peanuts, cocoanuts, olives, cotton-seeds, and corn.

Vegetable and animal oils and fats belong to the group of chemicals called ESTERS. Beef fat is largely *stearin.* Another ester is *palmitin.* Vegetable oils such as cotton-seed oil are mainly *olein.* Esters are widely distributed in nature and are of great economic importance.

Fill a small test-tube with *pure* olive oil. The label "pure olive oil" on can or bottle does not necessarily mean that it is pure olive oil. It has been found that a large portion of the "olive oil" sent to this country from Italy is actually corn oil flavored by treating with crushed olives from which most of the real olive oil has already been extracted.

Pure olive oil is necessary for this experiment. Set the tube containing the oil in a tumbler and pack it with cracked ice which has been sprinkled with salt. Soon a portion of the oil will solidify. Another portion will remain liquid. The solid portion is mainly PALMITIN. The liquid is OLEIN.

An Experiment with Camphor

CAMPHOR GUM

Camphor is the peculiar white resin of an evergreen tree with laurel-like leaves which grows in China and southern Japan. Synthetic camphor is manufactured from turpentine, borneol, and other substances by very complex chemical processes. Camphor is used in the manufacture of celluloid, explosives, and disinfectants.

Drop some small pieces of camphor into a dish filled with *clean* water. The camphor will move around in a lively manner over the surface of the water. Natural camphor is necessary for this experiment. The synthetic variety will not do.

If you touch the surface of the water with a spoon which has been greased with butter or olive oil, the motion of the camphor will cease. The thin film of oil or grease which spreads over the water will prevent it from acting on the camphor.

Camphor Solution

CAMPHOR GUM ALCOHOL

Drop a small piece of camphor into a test-tube and add some alcohol. Shake it up. The camphor dissolves rapidly. An alcoholic solution of camphor is known as spirits of camphor.

A Solvent for Turpentine

TURPENTINE ALCOHOL

Into a test-tube half filled with water, pour a teaspoonful of turpentine. Place your thumb over the mouth of the tube and shake well. The turpentine will be broken up into fine drops scattered through the water, but these drops soon collect together, forming a solid layer on top of the water. Turpentine and water will not mix. Water is not a solvent for turpentine.

Into another test-tube half filled with alcohol, pour a teaspoonful of turpentine. Place your thumb over the mouth of the tube and shake well. The turpentine dissolves and a clear homogeneous liquid results.

Turpentine is often used to remove paint spots from clothing. To remove the turpentine, use alcohol.

Making Ether

ETHYL ALCOHOL SULFURIC ACID

When a mixture of strong sulfuric acid and ethyl or grain alcohol is heated under the right conditions, the vapor which rises contains water and a volatile, inflammable liquid called ETHER. If the vapor is cooled and condensed, the liquids form two layers. The ether, being almost insoluble in the water and much lighter, forms the upper layer.

Ether is used as a solvent for fats and greases and in the manufacture of PYROXYLIN. Pyroxylin is used for making artificial leather, lacquers, and for a variety of other purposes. Pyroxylin lacquers and finishes are used on automobiles and are most familiar to the public under such trade names as Duco and Zapon.

When inhaled, ether produces insensibility to pain; hence it is used as an anesthetic in surgical operations. A mixture of ether and air explodes when ignited. Ether in small quantities in the atmosphere stimulates plant growth and is sometimes used to speed up or "force" fruit and flower growth.

Put 10 drops of undiluted sulfuric acid into a test-tube half-filled with grain or ethyl alcohol and heat the mixture gently over a low flame. Ethyl ether will be formed. It may be recognized by its characteristic odor. Some sulfur dioxide will also be released. Its disagreeable odor will be familiar if you have ever smelled the fumes from burning sulfur.

Only by using pure grain or ethyl alcohol (the chemist calls it ETHANOL) can you produce ethyl ether in this way. You can not purchase ethanol unless you are a chemist, physician, druggist, or licensed manufacturer. You need only a few spoonfuls. Perhaps your physician or druggist will give you that much. Denatured alcohol will not do.

When performing this experiment, the test-tube should be inclined at an angle in a clamp supported by the ring stand. Heat very gently, or the alcohol will be forced out of the tube and catch fire.

Black Dye from Tea

FERROUS SULFATE TEA

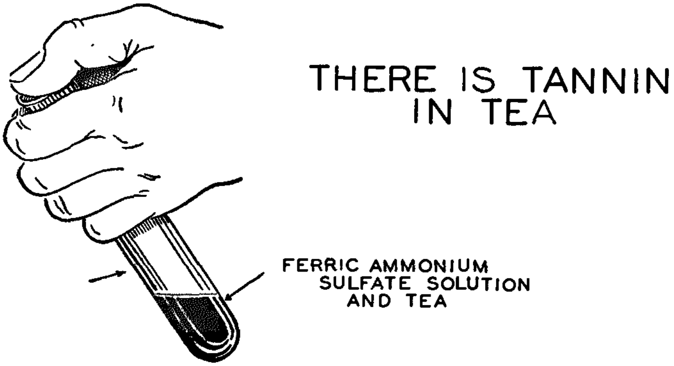
Dip a small piece of cotton cloth in some strong tea. Allow it to become dry; then dip it in a solution of ferrous sulfate and hang it up again.

Tea contains tannin, and black insoluble iron tannate will be formed in the fibers of the cloth.

Test to Show That There Is Tannic Acid in Tea

FERRIC AMMONIUM SULFATE TEA

Tea contains tannin, and its presence is very easily shown by the following test.



Iron and tannin form a black dye which is the basis of many writing inks. That there is tannin in tea is evident when an infusion of tea is mixed with a solution of ferric ammonium sulfate. The mixture turns black.

Pour some tea out of the tea-pot into a test-tube and then add a small pinch of ferric ammonium sulfate. Place your thumb over the mouth of the test-tube and shake the contents thoroughly. A black precipitate of iron tannate will form, showing that tea contains tannic acid.

Dyeing with a Mordant

LOGWOOD SODIUM CARBONATE ALUMINUM SULFATE

When cotton is dyed, the cloth is frequently first dipped in a MORDANT and then in the dye. A mordant is a substance capable of attaching itself both to the fibers of the cloth and, subsequently, to the dye also. Common mordants are the hydroxides of aluminum, tin, chromium, and iron. The name "mordant" is derived from the Latin *mordere,* meaning "to bite." Mordants are necessary in dyeing because certain dyes, used alone, will not attach themselves to the cloth.

Put 10 or 12 measures of logwood in a glass tumbler about ⅓ full of hot water. Logwood is obtained from a Central American tree. It is used as a dyestuff and may be obtained from a chemical supply-house. Let the solution stand for three-quarters of an hour to extract the color. Then add 10 measures of sodium carbonate and stir well with a glass rod.

Cut a strip of clean, white cotton cloth about 2 inches wide and 5 inches long. Wet half the strip with a solution consisting of 6 measures of aluminum sulfate dissolved in a test-tube full of water. Hang the strip of cloth up, and when it has dried, immerse it in the logwood and sodium carbonate solution. Let it soak in this dye for about 10 minutes.

When the cloth has been removed from the dye and dried, it will be seen that the portion which was treated with the solution of aluminum sulfate is darker than the other half. This dark portion is a "fast" color and will not "run" or "bleed" to as great an extent as the end not treated with the mordant. Frequently, the same dye will give different shades of color when used with various mordants.

CHAPTER IX

CHEMICAL TRICKS AND MAGIC

MANY professional magicians and entertainers use the changes wrought by chemical reactions in performing some of their tricks. The brilliant colors produced by mixing colorless solutions, the sudden appearance or disappearance of a substance, the amazing actions of many chemicals all seem like magic to any one who does not understand why these things happen. There are many experiments in this book which may be used to entertain your friends. Those given in this chapter are especially adapted for the exhibition of magic.

Changing a Copper Penny to "Silver"

MERCURIC NITRATE	COPPER PENNY

Mercury or quicksilver is the only metal which is liquid at ordinary temperatures. By ordinary temperatures is meant the temperatures at which we live. The old scientific name for mercury was *hydrargyrum*, from the Greek words meaning "water" and "silver." Mercury freezes at —40 degrees Fahrenheit, and boils at 662 degrees Fahrenheit. It is used for separating gold and silver from their ores, for coating mirrors, in barometers, as an expansive material in thermometers, in amalgams with other metals, and for many scientific purposes.

Clean a copper coin by rubbing it with metal polish. Then wash and dry it. Rub the coin with a small wad of cotton which is wet with a solution of mercuric nitrate. The mercuric nitrate solution can be prepared by dissolving 2 measures of mercuric nitrate crystals in 3 or 4 teaspoonfuls of water contained in a test-tube.

The copper coin will become coated with a film of metallic mercury and will appear like a silver rather than a copper coin.

A Magic Growth

MERCURIC NITRATE ALUMINUM

Aluminum oxidizes very readily and quickly. Fortunately, the thin film of oxide which forms on an aluminum surface protects the metal underneath from further oxidation. Otherwise, aluminum articles would soon crumble into a white powder.

You can produce a rapid and continuous oxidation of a piece of aluminum with the aid of mercuric nitrate. Using a knife blade, scrape a spot about as large as a quarter on the surface of a piece of aluminum. Then rub the spot with a small wad of cotton which has been dipped in a solution of mercuric nitrate. Set the aluminum aside and in a short time the spot which was thus treated will be covered with a layer of white powder. This is ALUMINUM OXIDE, and although you remove it innumerable times it will always grow again.

An interesting experiment is to scrape your initials on the outside of an aluminum pan and then daub them with mercuric nitrate solution.

Chemical Frost

EPSOM SALTS GLUE OR MUCILAGE

It used to be a custom at Christmas time to decorate the mirrors in barber shops with holiday greetings painted on the surface of the glass with imitation frost.

This is done with a saturated solution of magnesium sulfate or Epsom salts.

Heat a small beaker half filled with water. When the water is hot, add Epsom salts until no more will dissolve. This forms a saturated solution. Add a few drops of liquid glue, stir well, and apply this solution to a mirror or piece of glass with a wad of cloth or absorbent cotton. The liquid will begin to evaporate immediately and frostlike crystals will form on the surface of the glass.

Trick Matches

SODIUM SILICATE SOLUTION MATCHES

Paint the wooden sticks of some safety matches with sodium silicate solution. Do not get any of the silicate on the heads. Use a small camel's hair brush or hold the matches by their heads and dip the wooden sticks into the sodium silicate solution.

Lay the matches away until the coating of sodium silicate has dried. It will be invisible. Strike one of the matches. The head will flare up and go out. The stick will not burn. It has been made fire resistant.

Sodium silicate is used often to make woodwork, curtains, and other fabrics fire resistant.

Starting a Fire with Ice

AMMONIUM NITRATE ZINC DUST AMMONIUM CHLORIDE

Mix 4 teaspoonfuls of ammonium nitrate with 1 teaspoonful of ammonium chloride on a piece of paper. On top of this place 2 teaspoonfuls of zinc dust. Pour this mixture on a piece of melting ice. It will take fire and burn.

A Magic Bone

HYDROCHLORIC ACID CHICKEN BONE

Pour 6 test-tubes full of water into a beaker glass. Add 1 test-tube full of hydrochloric acid. Place a clean chicken-leg bone into the dilute acid and allow it to remain there for several days. If the bone is too long to lay flat in the beaker, cut it off to proper length with a saw.

The mineral part of the bone will be dissolved gradually by the acid. The flexible substance which is left will continue to retain the shape of the bone. It is called OSSEIN. When the ossein is dry, it has a translucent, horny appearance.

Making Gelatin from a Bone

The bones and skins of animals contain ossein. When ossein is boiled in water, gelatin is produced.

Place the flexible remains of the chicken bone left in the previous experiment in a pan filled with water and boil it gently for 3 or 4 hours. It will dissolve, and if there is not too much water present when the liquid cools, it will assume a jelly-like condition. Boiling the ossein produced GELATIN.

There is nothing "magic" about the slow process of making gelatin from a bone, but it is interesting from a chemical point of view. The experiment is suggested as a means of using the "rubber" bone produced in the previous experiment.

Changing Two Liquids into a Solid

CALCIUM CHLORIDE SULFURIC ACID

Put 4 teaspoonfuls of water into a test-tube. Add calcium chloride very slowly. Shake the solution from time to time and continue to add calcium chloride until no more will dissolve. This forms a saturated solution.

Add slowly ½ teaspoonful of sulfuric acid. The two transparent liquids will form a white, opaque solid.

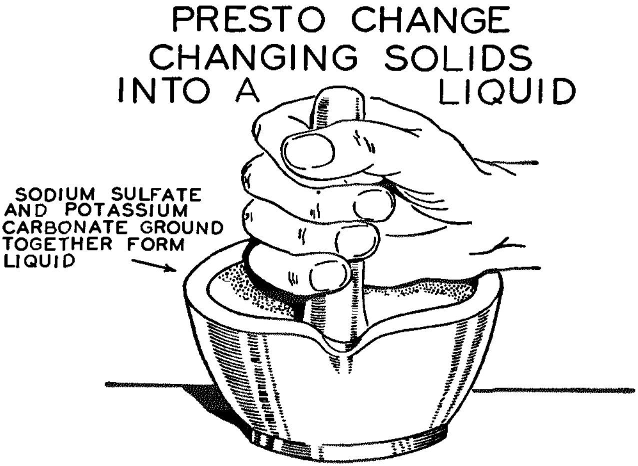
Presto Change: Two Solids into a Liquid

Mix 4 teaspoonfuls of sodium sulfate with 2 teaspoonfuls of potassium carbonate. Rub this mixture in a mortar and grind it. The two solids will form a liquid.

A "Rubber" Egg

Put a hen's egg into a tumbler and fill the latter with vinegar. Let it stand for 24 hours. If at the end of that time the shell is not soft and flexible, pour out the old vinegar and replace it with some which is fresh.

When the shell has finally become soft, you can push it into a milk bottle or through some other opening through which it would not ordinarily pass without breaking.



When sodium sulfate and potassium carbonate are ground together, the two solids become a liquid.

A Liquid from Solids

LEAD ACETATE SODIUM SULFATE

Place 1 teaspoonful of lead acetate in a test-tube. Add to this an equal amount of sodium sulfate. Shake the mixture and it will *turn to a liquid.*

The chemicals used in this experiment must be powdered. You can powder either lead acetate or sodium sulfate by rubbing in a mortar. Powder them separately, not together.

Chemical "Ice"

SODIUM SILICATE SOLUTION HYDROCHLORIC ACID

Put 3 or 4 teaspoonfuls of a thick sodium silicate solution in a test-tube. Add a small amount of hydrochloric acid and shake the tube. A gelatinous mass called SILICIC ACID is produced. If the solution of sodium silicate used is not too dilute, the silicic acid will adhere to the test-tube and have much the same appearance as ice.

Igniting Alcohol without a Match

ALCOHOL CHLORATE OF POTASH SULFURIC ACID

Put 2 measures of potassium chlorate into a beaker glass and add a teaspoonful of alcohol. Using a medicine dropper, sprinkle a few drops of sulfuric acid on the chlorate crystals. Do not hold your face above the beaker. Amidst the sputtering and crackling which takes place, there is so much heat evolved that the alcohol will be ignited.

Seventy-five years ago matches were made with chlorate of potash. The red mass on the end of the match stick consisted of chlorate of potash and sulfur, colored red by cinnabar (an ore of mercury). To ignite one of these matches, it was dipped into a "match flask." The match flask contained asbestos, moistened with sulfuric acid. The asbestos served to keep the acid from spilling and prevented the matches from being immersed too deeply.

A "Smoke" Trick

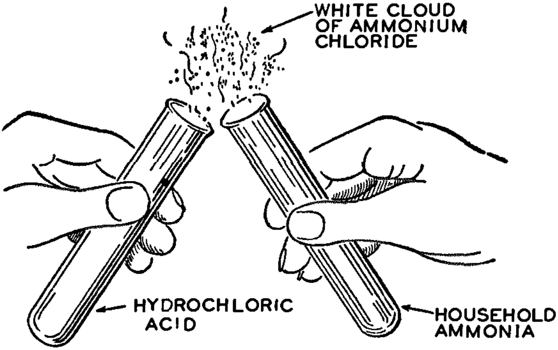
HYDROCHLORIC ACID HOUSEHOLD AMMONIA

Put a few drops of hydrochloric acid in a tumbler and spread it around with a glass stirring rod so that, as far as is possible, the inside of the glass is wet with acid. Pour out any excess acid.

Into a second tumbler pour some household ammonia. Rotate the tumbler so that its inside walls are thoroughly wet. Pour out any excess ammonia.

To an audience or to spectators who are not allowed to examine the tumblers, they will appear from a distance to be empty.

CHEMICAL SMOKE



The vapors of hydrochloric acid and ammonium chloride combine to form fine, smokelike white particles of ammonium chloride.

The tumblers, which up to this time must be kept far apart, should now be brought together and placed mouth to mouth. They will immediately appear to be full of smoke, and upon separating, the smoke will escape in a cloud.

The "smoke" consists of fine particles of ammonium chloride formed by the combination of the fumes of hydrochloric acid with ammonia. Do not delay the performance in the presence of your audience too long, or the ammonia will have escaped.

A "Smoke Screen"

AMMONIUM CHLORIDE

Heat some ammonium chloride on a piece of porcelain or in an old spoon over a flame. It will vaporize and give off thick clouds of smoke until it has entirely disappeared.

The Magic Thread

ALUM

Soak a piece of cotton thread about 2 feet long in a strong solution of alum. Then hang it up to dry on a hook or a nail that projects from the wall and away from any objects that might catch fire.

When the thread is dry, light the lower end and let it burn up. The ash will retain the form of the original thread and the thread will appear still to hang from the hook.

"Blowing" a Clear Liquid White

LIME-WATER

Fill a tumbler half-full of filtered lime-water and then blow into the liquid through a straw or a glass tube for a few minutes. The lime-water will become milky. The carbon dioxide in your breath reacts with the lime-water, or calcium hydroxide, as a chemist would call it, and forms calcium carbonate or chalk. Calcium carbonate is white and is insoluble in water. If undisturbed, the calcium carbonate will settle to the bottom after a time, leaving the water above quite clear.

A Red Liquid Which Can Be Blown Clear

SODIUM CARBONATE PHENOLPHTHALEIN SOLUTION

Put ½ measure of sodium carbonate into a tumbler half-filled with water. Add 4 drops of phenolphthalein solution and stir with a glass rod. The liquid will become red. Sodium carbonate has an alkaline reaction when dissolved in water and turns the phenolphthalein solution red. Phenolphthalein is red in an alkaline solution and colorless in an acid solution.

Blow through a straw or a glass tube into the red liquid contained in the tumbler. It will require a few minutes to produce results. Suddenly the liquid will become clear and look like ordinary water. Why did it change color? What happened?

The reason for this is simple to a chemist. Your breath contains carbon dioxide gas. Carbon dioxide dissolves in water and forms carbonic acid. By blowing into the water, you formed enough carbonic acid to neutralize the alkali causing the red color. Finally there was enough carbonic acid to make the water have an acid reaction. Result: the red color disappeared.

From Red to Green

ALCOHOL POTASSIUM DICHROMATE SULFURIC ACID

Dissolve 1 measure of potassium dichromate in water contained in a test-tube. The color of the solution will be orange-red. Add a few drops of sulfuric acid and heat the mixture slightly in a Bunsen flame. Then add some alcohol and the color will change from red to green.

Chameleon Mineral

MANGANESE DIOXIDE SODIUM HYDROXIDE

Grind together in a mortar 2 teaspoonfuls of manganese dioxide and 2 teaspoonfuls of sodium hydroxide.

Put this mixture in a porcelain crucible and heat it strongly in the flame of a Bunsen burner for 30 minutes.

Let the crucible and its contents cool. Then add some water to the black mass which it contains. A green solution will be formed. Put some of this solution in a test-tube and let it clear itself by settling for a time.

Add a few drops of sulfuric acid to the green solution in the test-tube and it will change immediately from green to red.

Why It is Called Chameleon Mineral

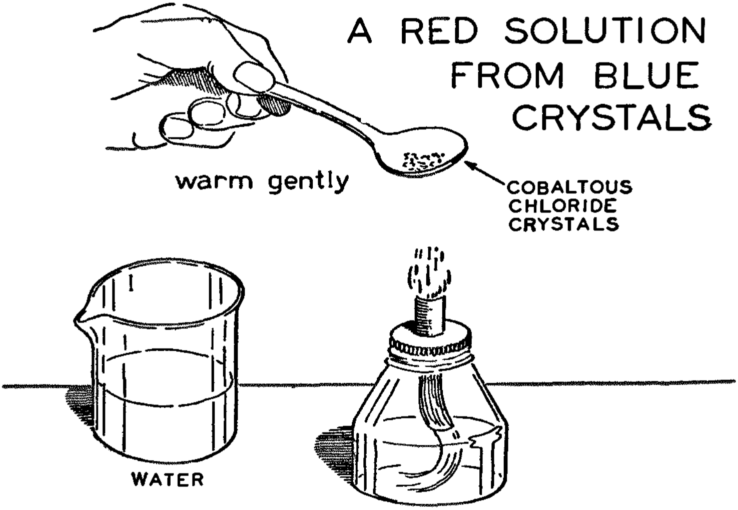
CHAMELEON MINERAL

Put some of the green solution of chameleon mineral in a beaker glass and leave it undisturbed. As the solution slowly absorbs carbon dioxide from the air, its color will undergo several changes. It will change to bottle-green, then violet, and finally will become crimson-red. A brown powder will be precipitated with the change in color.

Red Solution from Blue Crystals

COBALTOUS CHLORIDE

Place a few cobaltous chloride crystals in a small spoon and heat gently. Cobaltous chloride crystals normally contain a small amount of water and are ruby-red in color. When heated, the water is driven off and the crystals become deep blue.



The red crystals of cobaltous chloride contain water. When partially or completely dehydrated (deprived of water) they become deep-blue. Cobaltous chloride crystals may be dehydrated by warming gently. When dissolved in water, the blue crystals form a red solution.

Dissolve some of these blue crystals in water. They produce a reddish solution.

White, Red, and Blue

AMMONIUM HYDROXIDE PHENOLPHTHALEIN SOLUTION
COPPER SULFATE

Pour some ammonium hydroxide solution in a test-tube or beaker. Pure ammonium hydroxide is colorless. The household variety, since it contains impurities, is usually cloudy.

Add a few drops of phenolphthalein solution to the ammonium hydroxide. The mixture will become red. To this red solution add copper sulfate solution. It will change from red to deep blue.

Blue, Then White

SODIUM CARBONATE FERRIC AMMONIUM SULFATE
SODIUM FERROCYANIDE

Make up three solutions, as follows:

First, 2 measures of sodium ferrocyanide in a test-tube containing 2 teaspoonfuls of water.

Second, 2 measures of ferric ammonium sulfate in a test-tube containing 2 measures of water.

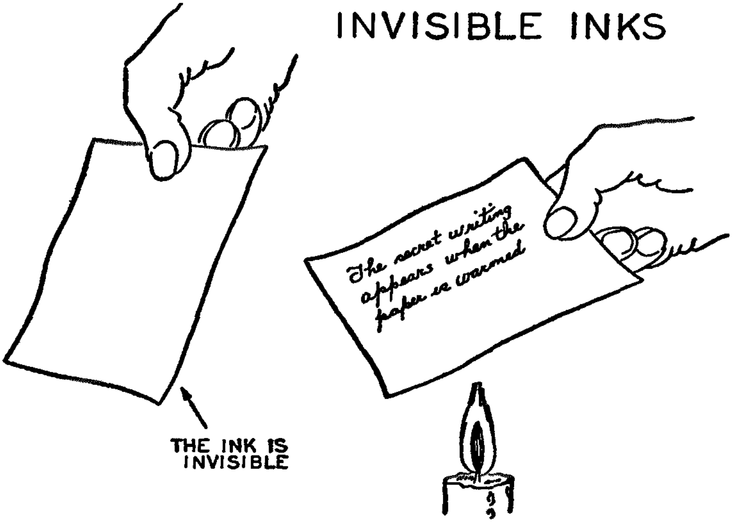
Third, 10 measures of sodium carbonate in a test-tube containing 2 teaspoonfuls of water.

Pour the solutions of sodium ferrocyanide and ferric ammonium sulfate into a small beaker. The mixture will be a beautiful blue. Add the sodium carbonate solution. The color will disappear.

Sympathetic Inks

LEMON JUICE MILK VINEGAR

Sympathetic inks are not real inks and of little practical value except as interesting experiments. Writing done with a sympathetic ink is invisible, or practically invisible, until developed by warming or heating. The writing should be done with a new, clean pen or it will not be invisible. Use clean white paper.



Writing done with "sympathetic" ink is invisible until it is revealed by heating or developing with the proper reagent. Lemon juice, grapefruit juice, vinegar, milk, onion juice, solutions of cobaltous chloride, ammonium chloride and ferric ammonium sulfate may be used as invisible inks. Writing done with these substances will not appear until heated.

Squeeze the juice out of a lemon and use it as an ink to write with a clean pen on white paper. When dry, the writing will be invisible but it will appear when the paper is heated over a Bunsen burner or alcohol lamp.

You can use milk as a sympathetic ink, and also vinegar.

Magic Writing Paper

TANNIC ACID FERRIC AMMONIUM SULFATE

Put 4 measures of ferric ammonium sulfate and 4 measures of tannic acid into a mortar and grind them together. Using a small wad of dry, white cotton cloth, rub this mixture all over the surface of a sheet of ordinary writing-paper lying on a flat surface.

Make sure that the mortar and pestle, your hands, the cloth, paper, and chemicals are perfectly dry. The slightest trace of moisture will produce black smudges and spoil the experiment.

Dust off any surplus powder remaining on the paper. Then, using a new clean pen-point dipped into water instead of ink, write on the prepared surface. The writing will be black just as if you had dipped the pen into ink.

Magic with Tea

FERRIC AMMONIUM SULFATE

Dissolve 2 measures of ferric ammonium sulfate in 2 teaspoonfuls of water. Using a clean pen dipped into this solution, print or draw a rough sketch on a sheet of white writing-paper.

When the sketch or printing is thoroughly dry it can be made to appear by wetting the paper with strong tea.

A Trick with Salt

SODIUM CHLORIDE

Put 2 teaspoonfuls of water into a test-tube and add common table salt (sodium chloride) until no more will dissolve. This will form a saturated solution.

Write with this solution on white paper, using a clean pen. When the writing is dry, rub it with the under side of the bowl of one of your aluminum measuring spoons. The writing will then become visible.

Writing Disappears by Magic

SODIUM CARBONATE SODIUM FERROCYANIDE
FERRIC AMMONIUM SULFATE

Add 2 measures of sodium ferrocyanide and 2 measures of ferric ammonium sulfate to 2 teaspoonfuls of water contained in a test-tube. Shake the mixture until both salts have dissolved. This will form a blue-colored ink. Using a clean pen, write with some of this ink on a sheet of white paper.

When the ink has dried, dip a wad of cotton into a solution prepared by dissolving 8 measures of sodium carbonate in 2 teaspoonfuls of water. Swab the writing with the cotton wad which is wet with the sodium carbonate solution and it will disappear.

CHAPTER X

MISCELLANEOUS EXPERIMENTS

How to Distill Water

PURE water is an important material in the chemical laboratory.

Water is an almost universal solvent. It is the remarkable ability of water to dissolve nearly all other known substances which makes it so useful. Many chemical reactions can not take place except in a solution. Many substances which have no effect upon one another while dry will act upon one another when dissolved in water.

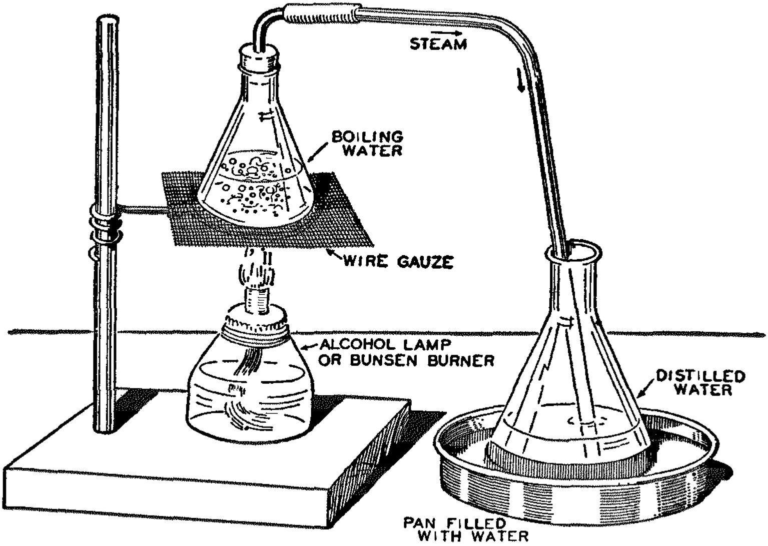
Almost all water, except that which has been *distilled,* contains invisible chemicals in solution. The process of distilling consists of boiling a liquid until it vaporizes, and then condensing the resulting vapor. The distillation of water produces the purest water that can possibly be obtained in any way.

A one-quart bottle of distilled water can be purchased at an automobile-supply store for 10 or 15 cents. You can distill it yourself if you wish.

Put some water into a flask equipped with a glass delivery tube which leads into another flask as shown in the illustration. The second flask rests in a shallow pan of cold water.

A good method of aiding the condensation is to wrap a wet cloth around the receiving flask and keep it wet with cold water.

DISTILLING WATER



In order to get pure water, it must be distilled. Distilling water consists of boiling it and condensing the steam which is formed. To cool and condense the steam, the chemist generally uses a condenser. A condenser is a glass tube kept cool by surrounding it with cold water. A simple apparatus for distilling water is illustrated.

The flask containing the water to be distilled is placed on a ring stand over an alcohol lamp or Bunsen burner. When the water boils, steam passes through the delivery tube into the condensing flask. The distilled water that results from this method of purification is called the *distillate.*

Hard Water

SOAP LIME-WATER

Calcium and magnesium occur in the soil and in many minerals and rocks. The water of many springs, rivers, and wells contains these two elements.

Water containing salts of lime or magnesia in solution is known as *hard* water. Soap will not produce a lather with hard water until enough soap has been dissolved in the water to precipitate the calcium in the form of calcium compounds of oleic, stearic, and palmitic acids.

When ordinary soap is dissolved in hard water, a white flocculent precipitate is formed. Special soaps called "hard-water soaps" are made for use with water containing lime.

Dissolve a little soap in hot water and add some clear lime-water to it. The soap solution will become turbid. White flakes which feel sticky to the fingers will be deposited on the sides of the vessel.

To Make Lime-Water

LIME

Put 1 teaspoonful of lime into a clean milk bottle or one-quart fruit jar filled with water. Shake it up thoroughly and then let it settle. Pour off the clear liquid at the top into another bottle and keep it tightly corked. This clear liquid is LIME-WATER. It contains a small amount of CALCIUM HYDROXIDE in solution and is useful in testing for carbon dioxide.

Heat by Chemical Reaction

SODIUM HYDROXIDE ALUMINUM POWDER

The reaction which takes place in this experiment has already been described for the purpose of making hydrogen. The gas which is liberated is HYDROGEN.

Place 1 heaping teaspoonful of sodium hydroxide and an equal amount of aluminum powder in a small beaker. Add enough water to cover the mixture. It will immediately become quite hot. Heat has been liberated by the chemical reaction which takes place.

Some chemical reactions liberate heat. Others absorb heat. EXOTHERMAL ACTION, as those reactions which produce heat are called, are usually spontaneous. ENDOTHERMAL ACTIONS are those which absorb heat. Usually, they are not spontaneous.

Heat and Cold from a Solution

SODIUM SULFATE

Put several crystals of sodium sulfate (Glauber's Salt) in an evaporating dish and heat them very gently. Heating them in a sand bath or on a wire gauze well above the flame is best. The crystals will soon become covered with an opaque white coating and will finally crumble into a white powder. The white powder is ANHYDROUS SODIUM SULFATE.

Put some water into a beaker and immerse a thermometer in the water. What is the temperature of the water? Put 2 teaspoonfuls of *ordinary* sodium sulfate into the water and stir. Take the temperature of the solution. It is much COLDER.

Put some water into another beaker and take its temperature. Dissolve 2 teaspoonfuls of *anhydrous* sodium sulfate in the water. Stir the mixture and take its temperature. It is much WARMER than it was before.

Freezing by Evaporation

CARBON DISULFIDE

When a dish of water is left open to the air for several days, all or most of the water moves out of the dish into the surrounding air. EVAPORATION, a change from the liquid to the gaseous state, has occurred.

Science explains this in the following manner. The molecules which make up the water in the dish are continually moving back and forth in every direction. Some of the molecules near the surface, which happen to be moving in the right direction, jump away from the water and move about in the air immediately above the surface of the water. Then, they are no longer in the liquid state but are gaseous molecules, or in other words WATER VAPOR.

When a molecule changes from the liquid state to vapor it absorbs heat; it withdraws this heat from the liquid which it left. If evaporation takes place rapidly, there is a noticeable cooling of the liquid. Some liquids evaporate more rapidly than others at ordinary temperatures. In fact, they evaporate so rapidly that their own temperature may drop below the freezing point of water (32 degrees Fahrenheit).

Put 2 teaspoonfuls of carbon disulfide into a small beaker or a watch-glass. Set the glass upon a wet piece of wood. By means of a tire pump or by blowing through a glass tube, direct a current of air across the surface of the liquid.

Carbon disulfide is extremely volatile, that is, it evaporates rapidly. In so doing it becomes so cool that the wet beaker or watch-glass is frozen to the wood.

This experiment should be performed where there is a good draught of air and away from any flames. Carbon disulfide is more inflammable than gasoline, has an unpleasant odor, and its fumes are somewhat poisonous if inhaled in quantity.

How to Make Potassium Manganate

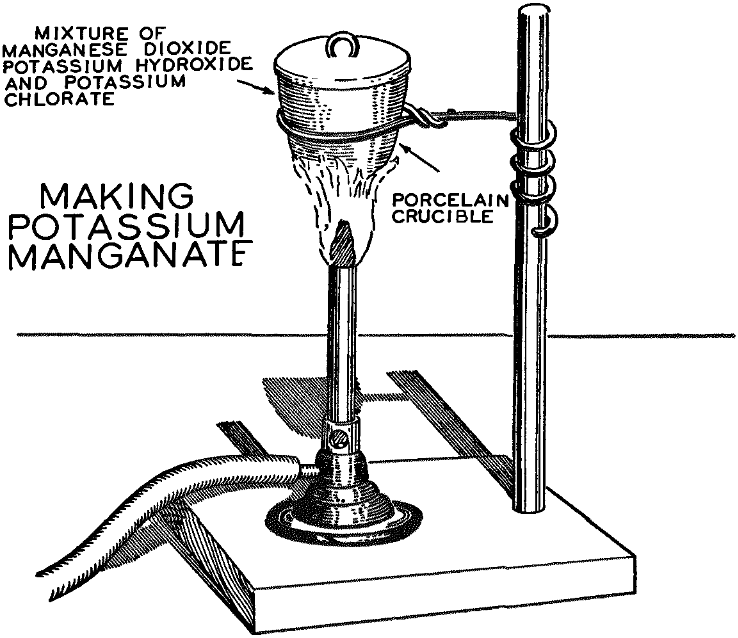
MANGANESE DIOXIDE POTASSIUM HYDROXIDE
POTASSIUM CHLORATE

In a small porcelain crucible put 1 teaspoonful of powdered manganese dioxide, 1 teaspoonful of potassium hydroxide (caustic soda), and ½ teaspoonful of potassium chlorate. Heat this in the flame of a Bunsen burner until the mixture turns dark green. The green substance is POTASSIUM MANGANATE, a compound of potassium, manganese, and oxygen. Dissolve some of the potassium manganate in water. It will form a dark-green solution.

The Preparation of Potassium Permanganate

POTASSIUM MANGANATE WATER

Half fill a clean test-tube with water. Add some potassium manganate. When it has dissolved, heat the solution until it



This experiment shows how a metal sometimes behaves like an acid. All metals combine with acids to form the class of chemicals called salts. The sulfates, nitrates, chlorides, carbonates, etc., are salts.

Many metals also have the ability when in combination with oxygen to act like acids and combine with other metals. Manganese and chromium have this property. Manganese and oxygen acting together like an acid form manganites, manganates, and permanganates. Chromium and oxygen form chromites, chromates, and dichromates.

boils. The color will change from green to purple. The purple color is due to the formation of POTASSIUM PERMANGANATE.

When a solution of potassium permanganate is evaporated, it forms dark purple, almost black, crystals having a greenish luster.

Potassium permanganate is a valuable disinfectant. It is also a bleaching agent. It is useful to the chemist in some of his analytical processes. A solution of potassium permanganate is used sometimes for dyeing leather and for coloring wood a deep brown. A *dilute* solution may be used for allaying the itching produced by poison ivy.

Oxidization by Means of Potassium Permanganate

POTASSIUM PERMANGANATE FERROUS SULFATE
SULFURIC ACID

Potassium permanganate gives up its oxygen very readily and changes to a HYDROXIDE OF MANGANESE. In order to accomplish this, it is merely necessary to add an oxidizable substance to a solution of potassium permanganate. If the solution is an acid solution, the hydroxide dissolves, forming a colorless solution.

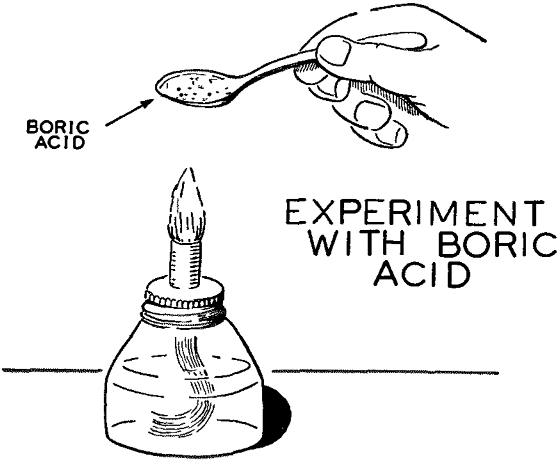
Put 2 measures of ferrous sulfate into a test-tube containing 3 teaspoonfuls of water. Add 3 drops of sulfuric acid. When the ferrous sulfate has dissolved, add, drop by drop, a dilute solution of potassium permanganate. The color of the potassium permanganate will be destroyed as it gives up oxygen to the ferrous sulfate.

An Experiment with Boric Acid

BORIC ACID

Although the chemical element called boron is a member of the aluminum family, it is not a metal. Boron and its compounds greatly resemble carbon and their compounds in their chemical properties.

Boron is familiar chiefly in the form of its acid, known as *boric* acid, and in the form of borates. Borates are salts of boric acid. The most common borate is BORAX or, as the chemist calls it, SODIUM TETRABORATE.



Not all acids are liquids. Some, like boric, citric, oxalic, tannic, and salicylic acids are solids. Boric acid is obtained in beautiful lustrous scales from a native calcium borate found in California. Boric acid is a very weak acid. When dissolved in water it has antiseptic properties.

When heated, the commercial form of boric acid will dissolve in its own water of crystallization. After all the water has been driven off, anhydrous boric acid remains.

Under the name of *boracic acid,* boric acid is used in medicine as an antiseptic solution. It is especially valuable as an eye-wash.

Heat some boric acid in an old teaspoon or tablespoon. The acid will melt first and dissolve in its water of crystallization. As the heat is continued, the water will become steam, the molten mass will become pasty and will swell up as the steam escapes.

After all the water has been driven off, a clear viscous liquid will remain. This is ANHYDROUS BORIC ACID. *Anhydrous* means "without water."

Touch the surface of the liquid with a glass rod or a small stick. You will find, if you pull the stick away slowly, that you can draw out long threads of boric acid.

If the fused mass in the spoon is allowed to cool, it will contract and crack in every direction.

How to Make Boric Acid

SODIUM BISULFATE BORAX

Put 4 teaspoonfuls of borax into a small beaker half filled with water. Heat the beaker slightly and stir with a glass rod until all the borax has dissolved.

Put 2½ teaspoonfuls of sodium bisulfate into a second small beaker half filled with water. Heat this if necessary to dissolve the solid. Then pour this solution into the borax solution and set the mixture away to cool.

After a while, white crystals will form in the solution. These crystals are boric acid.

Boric acid (boracic acid) is found in jets of water vapor which issue from the ground in the volcanic region of Tuscany, Italy. Brick-work basins placed over the openings condense the vapor, and from this water boric acid is obtained in the form of the same crystals which are produced in this experiment.

How to Make Potash

WOOD ASHES

Potassium is a silver-white metal which unites with oxygen so violently that it must be kept under petroleum. Since petroleum does not contain oxygen itself nor dissolve oxygen from the air to any extent, it acts as a shield for the potassium.

Through experiment, chemists have learned that plants need six major foods in order to be healthy and to grow to maturity. These foods are potassium, phosphorus, sulfur, nitrogen, calcium, and magnesium. In addition, minute quantities of zinc, iron, manganese, and boron are necessary.

Back in pioneer days, when soap was made at home, the potassium used for this purpose was secured from wood ashes in the form of POTASH.

Put ½ cupful of wood ashes into a small enamelware pan and pour 1 pint of boiling water on the ashes. Stir the mixture well.

Test some of the mixture with litmus-paper. It will turn red litmus-paper blue, showing that it is alkaline.

Separate the liquid from the solid matter in the mixture by filtering. Put a few drops of the clear liquid into a test-tube and add 2 or 3 drops of dilute acid. You may use sulfuric or hydrochloric acid. There will be a brisk effervescence, and from this you may safely surmise that the liquid contains a carbonate.

Evaporate the solution to dryness in a porcelain dish. The salt which is obtained is crude potash or POTASSIUM CARBONATE. Refined potash is known as pearlash.

A Deflagration

POTASSIUM NITRATE CHARCOAL

Potassium nitrate is also known as saltpeter. It is white and has a cooling, bitter taste. Corned beef is produced by soaking fresh beef in a solution of saltpeter and common salt. Potassium nitrate is used in the manufacture of gunpowder and of fireworks.

Mix a teaspoonful of powdered saltpeter and an equal quantity of powdered charcoal thoroughly together. Put the mixture in a heap on a brick and ignite it with a red-hot wire.

When the deflagration (chemist's term for a sudden, sharp combustion) is over, a white solid material will be found on the brick. Scrape some of this up and dissolve it in water contained in a test-tube. Add a few drops of hydrochloric acid. The brisk effervescence is the escape of carbon dioxide.

When the deflagration took place, the potassium nitrate oxidized the carbon, forming carbon dioxide. Some of the carbon dioxide entered into combination with the potassium, forming POTASSIUM CARBONATE.

Fire Writing

POTASSIUM NITRATE

Add 1 teaspoonful of potassium nitrate to 2 teaspoonfuls of water contained in a test-tube. Using an artist's camel's-hair brush as a pen and this *strong* potassium nitrate solution as the ink, write or print on a sheet of unglazed paper. Make

Potassium nitrate or *saltpeter* gives up part of its oxygen when heated. Paper which has been soaked in potassium nitrate solution and dried becomes *touch* paper. It will burn slowly without flame.

A strong solution of saltpeter used as an ink to write upon paper is the secret of the trick called fire-writing. When dry, the writing is not visible until touched with a glowing splinter.

the writing large and bold. When it has dried, apply the spark on a piece of smoldering cotton cord or on the end of a glowing splinter to the writing. It will ignite and burn like a fuse.

The Preparation of Silver Nitrate

SILVER COIN NITRIC ACID

Place a silver coin in a beaker. Add a small amount of water and then pour in a little nitric acid. The acid will attack the silver and begin to dissolve it. If the action stops, add a little more acid, but do not use any more acid than is necessary in order to dissolve the silver completely.

The solution in the flask will be blue. It is a solution of cupric nitrate and silver nitrate. All silver coins contain a certain amount of copper, used for the purpose of hardening the silver. It is the copper in the solution which produces the blue color.

When the silver has dissolved, heat the solution gently so as to evaporate the liquid and leave a mixture of silver nitrate and cupric nitrate crystals. Put this solid residue in a porcelain crucible and heat over a Bunsen flame to a temperature of 300 degrees Fahrenheit. At this temperature, the silver nitrate is unaffected, but the cupric nitrate is broken down into CUPRIC OXIDE.

Cupric oxide will not dissolve in water but silver nitrate will. When the mixture in the crucible has cooled, mix it with water so as to dissolve the silver nitrate and pour into a filter. The cupric oxide will remain on the filter paper, but the silver nitrate will pass through in solution or as a filtrate. If this solution is evaporated carefully, crystals of silver nitrate will remain.

Silver nitrate, either in solution or in crystal form, should be kept in a bottle of dark glass or in one wrapped in dark paper. It is affected by light.

Light and Silver Nitrate

SILVER NITRATE

Silver nitrate is used sometimes by doctors for cauterizing or searing wounds, infections, etc.

Silver-nitrate crystals should not be touched with the fingers but always handled with a small spoon or a pair of tweezers. The alchemists used the Latin word *luna,* meaning "moon," as the name for silver. Thin sticks of silver nitrate are called *lunar caustic* in the drug trade.

A solution of silver nitrate, as well as the crystals, should be handled with care so as not to spill it or get it on the fingers.

Pure silver nitrate is not affected by light, but when it is deposited upon cloth or upon the skin it turns black upon exposure to light. It is quite indelible and for that reason is an ingredient of marking inks.

Make a solution of silver nitrate. Using a small glass stirring rod as a pen and the silver nitrate solution as ink, mark on a piece of white cotton cloth. Expose the cloth to the sunlight. It will turn black wherever it was wet with the nitrate solution.

Making Silver Chloride

SILVER NITRATE SODIUM CHLORIDE

Dissolve a few small crystals of silver nitrate in a clean test-tube about one-third filled with water. Add a few drops of a solution of common salt. The insoluble, curdy, white precipitate which forms is composed of silver and chlorine. It is SILVER CHLORIDE.

Like some other compounds of silver, silver chloride darkens on exposure to light. First it turns violet, then brown. In solution, and unseen until the liquid is evaporated, is SODIUM NITRATE.

An Experiment with Silver Chloride

SILVER NITRATE
NITRIC ACID
SODIUM CHLORIDE
AMMONIUM HYDROXIDE

Make some fresh silver chloride by adding a solution of common salt to some silver nitrate solution. Pour this into a filter so that the liquid drains off and the silver chloride remains. Wash the chloride by squirting water on it from a wash bottle while it is still on the filter.

Put some of the washed silver chloride in a test-tube and add some nitric acid. You will find that it will not dissolve in nitric acid.

Put another portion of the washed silver chloride in a test-tube and add some strong household ammonia. It will dissolve.

To Separate the Silver and Copper in a Coin

SILVER COIN
SODIUM CHLORIDE
ZINC
NITRIC ACID
SULFURIC ACID

Dissolve a silver coin in dilute nitric acid. This process has been described in a previous experiment.

When the coin has disappeared, add enough water to the solution to make ½ pint. Then add a strong solution of common salt until there ceases to be any further precipitate.

This white silver chloride which is formed thus should be filtered. While it is still on the filter, wash it thoroughly, using hot water in the wash bottle. Then put the funnel containing the filter paper and silver chloride in a warm place where it will dry.

When the silver chloride is dry, put it into a porcelain crucible and heat gently until the chloride melts. Then drop in on top of the melted chloride a piece of clean sheet-zinc, approximately large enough to cover the bottom of the crucible.

Next add 3 or 4 teaspoonfuls of water and a few drops of dilute sulfuric acid. Let the crucible and its contents stand for at least 24 hours.

At the end of that time, the silver chloride will have been reduced to silver, and zinc chloride formed. Take out the zinc and put the remaining contents of the crucible into a filter. Wash the metallic silver which remains on the filter paper with dilute sulfuric acid and then with water.

If you heat the metallic silver on charcoal in the blowpipe flame, it will melt and form a globule of almost pure silver. It has been separated from the copper with which it was alloyed in the coin.

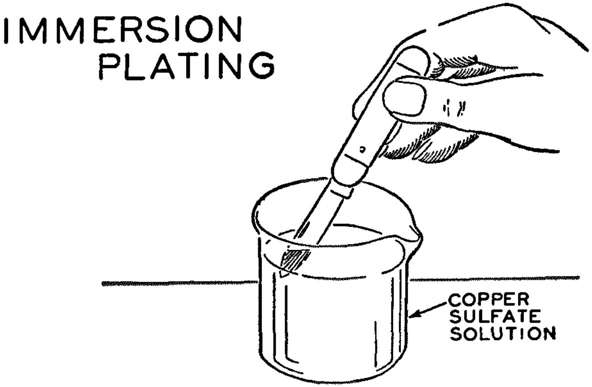
Silver and Lead Iodide

SILVER NITRATE LEAD NITRATE POTASSIUM IODIDE

Drop a small crystal of silver nitrate into a test-tube containing 2 or 3 teaspoonfuls of water. Shake the tube until the crystal dissolves.

To a second test-tube containing about the same amount of water add a measure of lead nitrate.

To each tube add a few drops of a solution of potassium iodide. Yellow will be produced in the first tube and orange in the second.



The steel blade of a pocket-knife or screw-driver can be coated with copper by dipping it into a solution of blue vitriol or copper sulfate. The same principle is used to provide sewing needles with gold-plated eyes. They are dipped in gold chloride. This method of coating one metal with another is called immersion plating and is a similar chemical action to the immersion deposition which takes place in The Silver Tree Experiment.

The Silver Tree

SILVER NITRATE

COPPER WIRE

Put a few small crystals of silver nitrate into a 1-ounce bottle not quite filled with distilled water. When the crystals have dissolved, push a short piece of clean copper wire into the under side of the cork and put it in the bottle. The wire should dip into the silver nitrate solution.

Put the bottle where it will not be disturbed. A gleaming pendant of crystals of pure silver will grow down from the copper wire.

This process is called the replacement of metals. Some of the copper dissolves in the solution, becoming COPPER NITRATE.

Some Properties of Iodine

TINCTURE OF IODINE CARBON TETRACHLORIDE

Iodine is a solid, almost black in color. When heated, it vaporizes and gives off a violet vapor. Iodine dissolved in alcohol is called tincture of iodine and forms a dark-brown solution. When dissolved in carbon disulfide or chloroform, iodine forms a violet-colored solution.

Pour enough carbon tetrachloride into a test-tube to half-fill it. Add a drop of tincture of iodine. The resulting solution will be pink.

The Violet Vapor of Iodine

IODINE

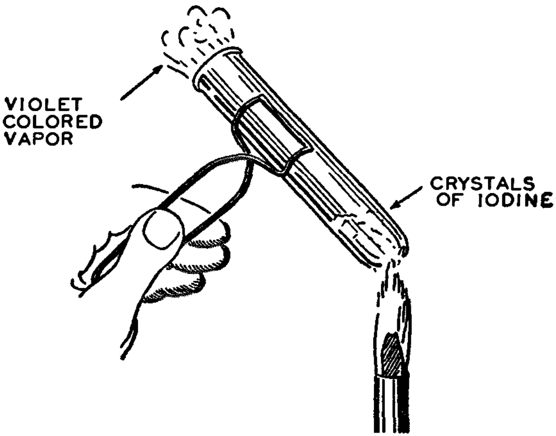
Iodine is a member of a family of chemicals called the HALOGENS which includes chlorine, bromine, and fluorine.

Iodine is most familiar in the form of the alcoholic solution called tincture of iodine, used as an antiseptic for surface cuts and scratches.

Pure iodine is a solid, almost black in color. Iodine may be easily SUBLIMATED or, in other words, converted from a solid to a vapor by heat and then solidified again by cooling.

Put a small crystal of iodine into a dry test-tube. Heat the test-tube gently in a flame and notice that the iodine changes into vapor without first changing into a liquid. The vapor is violet-colored and very heavy. If the upper part of the tube is cold, the vapor will condense and form minute iodine crystals there.

THE VIOLET VAPOR OF IODINE



The elements iodine, chlorine, bromine, and fluorine form a chemical family called the halogens or "salt makers." Fluorine and chlorine are gases. Bromine is a dark-red liquid and iodine is a solid. Iodine is a solid substance but is easily changed to vapor as demonstrated in the experiment illustrated above. At first, the vapor has a reddish-violet tint. If strongly heated, it becomes deep blue.

Iodine Dissolves in Alcohol

ALCOHOL **IODINE**

Drop a very small piece of iodine into a test-tube half-filled with alcohol. The alcohol will dissolve the iodine and a brown solution will be formed. This is TINCTURE OF IODINE.

Iodine Forms a Violet Solution in Carbon Disulfide

CARBON DISULFIDE IODINE

Iodine is quite soluble in carbon disulfide and in chloroform or ether.

Put a very small piece of iodine into a test-tube and add a few drops of carbon disulfide. The iodine will dissolve and form a *violet*-colored solution.

The Solubility of Iodine

IODINE CARBON DISULFIDE

If we put a small particle of iodine into a test-tube half-filled with water, we will find that it dissolves very slowly. The solution will have a scarcely perceptible brown tint. Eventually it will be saturated, that is, contain all the iodine it can, but it will still be a very dilute solution, for at ordinary temperatures it requires about 700 gallons of water to dissolve 1 pound of iodine.

Add a teaspoonful of carbon disulfide to the water and iodine in the test-tube and shake it up well. The greater part of the iodine will find its way into the carbon disulfide. On standing for a few moments, the carbon disulfide will settle to the bottom of the tube and it will be a reddish-violet color.

Carbon disulfide is used by chemists to discover the presence of iodine in water and it will detect the presence of 1 part of iodine in 1,000,000 parts of water. The process of withdrawing a substance from one solution and securing it in another is called EXTRACTION.

An Experiment with Potassium Iodide

POTASSIUM IODIDE IODINE

Put 1 measure of potassium iodide into a clean test-tube half filled with water. When the potassium iodide has dissolved, drop a piece of iodine into the solution and shake it up. The iodine will dissolve. A water solution of potassium iodide, hydrogen iodide, or any other iodide will dissolve large quantities of iodine.

Smelling Salts

AMMONIUM CHLORIDE CALCIUM OXIDE

Mix ½ teaspoonful of calcium oxide (quicklime) with ½ teaspoonful of ammonium chloride (sal ammoniac). Place this mixture in a test-tube and warm it for a few minutes over a flame. A gas will be given off. Smell it cautiously. This gas (ammonia) has the property of reviving people who have fainted or who feel faint, and is therefore used in "smelling salts." One form of smelling salts consists of ammonium carbonate crystals mixed with lavender perfume. Ammonia gas "flavored" with the scent of lavender is given forth.

How to Make Ammonium Sesquicarbonate

AMMONIUM CHLORIDE CALCIUM CARBONATE

Mix 1 teaspoonful of ammonium chloride (sal ammoniac) with ½ teaspoonful of calcium carbonate. Put the mixture in a 4-ounce Erlenmeyer flask and place it in a sand bath over a Bunsen burner.

As soon as a pungent odor is observed to be issuing from the flask, invert a larger flask over it. The fumes will condense in the upper flask and form AMMONIUM SESQUICARBONATE. This is a white crystalline salt which has a faint odor of ammonia.

Borax as a Flux

SHEET-COPPER
IRON WIRE
BORAX
BRASS FILINGS

Borax, or sodium tetraborate as it is known in chemistry, is used as a *flux* in hard soldering. A flux is a substance used to bring infusible bodies into a fusible form of combination.

Place a piece of iron wire about 1 inch long on a piece of sheet-copper the size of a quarter. The copper should rest on a block of charcoal. Sprinkle some coarse brass filings over the wire, and then on top of the brass filings sprinkle some borax. Direct the flame of a blowpipe on the iron wire, brass, and borax. It may be necessary to add some more brass and borax. If heated to a bright red, the brass will melt and BRAZE the iron wire to the copper.

Before heating, the iron and the copper should be brightened by rubbing with sandpaper or scratching with a file.

Ammonium Chloride as a Flux

AMMONIUM CHLORIDE
SOLDER
SHEET-COPPER

Heat a piece of clean sheet-copper in the flame of a Bunsen burner and drop a piece of solder on it. The solder will melt but it will not unite with the copper firmly. When it has cooled you can knock or pry the solder loose from the copper.

Repeat the experiment. This time, before you put the solder on the hot copper, sprinkle some ammonium chloride (sal ammoniac) on the copper. Then add the solder. The solder will melt and unite firmly with the copper. This time you will not be able to pry the solder loose. Sal ammoniac is a flux, frequently used for soldering. Heat decomposes the sal ammoniac. The hydrogen chloride which is released removes the layer of oxide on the copper. Unless the copper oxide is removed, it prevents the solder from uniting with the copper.

An Experiment with Phosphorus

RED PHOSPHORUS

There are two kinds of phosphorus.

Ordinary or yellow phosphorus is a waxlike solid which is exceedingly poisonous. It is necessary to keep it under water to avoid spontaneous combustion.

Red phosphorus is a scarlet-red color, has no odor, is not poisonous, and does not need to be kept under water. Red phosphorus is obtained by heating yellow phosphorus to 450 degrees Fahrenheit in a vessel from which all air has been excluded.

Yellow phosphorus is dangerous for any one except a trained chemist to experiment with. Red phosphorus can be safely handled by the student or young experimenter.

Close one end of a piece of glass tubing about ¼ inch in diameter and 12 inches long by holding it in a Bunsen flame. Turn the tube from time to time while heating it. When the end is closed, do not lay it down until it has cooled or it may crack.

When the tube has cooled, put a measure full of *red* phosphorus into the tube. Heat the phosphorus gently above a flame. It will vaporize and condense on the cooler portions of the tube. It will no longer be red, but will be yellow. The red phosphorus has been changed back into ordinary or yellow phosphorus. The amount of yellow phosphorus produced thus is too small to be harmful.

The reason for using a long piece of tubing in this experiment, in place of a test-tube, is to exclude air. The air in the narrow tube is deprived of its oxygen by the combustion of a small portion of the red phosphorus. The remaining phosphorus is thus protected from further oxidation.

The striking surface on a box or package of safety matches is a thin layer of antimony trisulphide mixed with red phosphorus and glue.

To Make Steel Out of Iron

POTASSIUM FERROCYANIDE IRON NAIL

Heat an iron nail red-hot and dip it into some powdered potassium ferrocyanide. If some of the ferrocyanide clings to the nail, do not brush it off. Heat the nail red-hot again and quickly plunge it into cold water.

The surface of the nail where it was in contact with the potassium ferrocyanide will prove to be so hard that it resists the action of a file. Carbon from the potassium ferrocyanide was absorbed by the iron nail and formed a coating of steel on its surface.

An Experiment with Calcium Sulfate

PLASTER OF PARIS

Gypsum, or calcium sulfate, is, next to the carbonate, the most widely distributed and the most abundant salt of calcium. When heated to 350 degrees Fahrenheit, gypsum loses three-quarters of its water of crystallization and becomes the chalky powder known as plaster of Paris.

When water is mixed with plaster of Paris, a reaction occurs which is the reverse of that used in its manufacture. A hardening, known as "setting," takes place.

Plaster of Paris is used in making casts, molds, and in surgery. When it is mixed with water it sets in a short time and forms a rigid mass. The setting of the plaster is accompanied by a slight increase in volume. This makes a very sharp reproduction of all the details of the object from which a cast is made.

Wind a strip of paper around the edge of a medal or souvenir pocket piece so as to make a shallow cylindrical box, the bottom of which is formed by the medal. Secure the loose end of the strip of paper with paste or other adhesive.

Mix 2 teaspoonfuls of plaster of Paris with enough water to make a paste and pour it into the box. In half an hour it will become so hard that both the medal and the paper can be removed. A reversed impression of the medal will appear on the cast.

After this has dried for 24 hours it may be used as a mold. Smear the impression with oil and wind a strip of paper around the edge so as to form a cylindrical box with the impression as the bottom of the box. Mix some more plaster of Paris with water and pour it into the box. When this has hardened it may be removed from the mold. It will be a true or positive impression of the medal.

The "ivory" surface of decorative plaster casts, which makes them washable, is secured by painting with a solution of paraffin dissolved in gasoline. This fills the pores of the plaster and prevents disintegration of the plaster by moisture and water.

Bismuth, a Metal Which Burns with a Blue Flame

BISMUTH

Heat a small piece of bismuth on charcoal by means of a blowpipe. At ordinary temperatures bismuth remains unchanged, but when heated to redness it burns with a bluish flame. The yellow powder which is formed is an OXIDE OF BISMUTH.

Bismuth is a shining, brittle, grayish-white metal with a reddish tinge. It looks much like antimony but can be distinguished by its red tint.

Bismuth is neither rare nor plentiful. American bismuth is obtained as a by-product in the refining of lead and copper. Unlike other metals which contract when they cool, bismuth has the property of expanding when changing from the liquid to the solid state. This makes it valuable as an ingredient of type metal, where it aids in filling the mold and producing an accurate impression.

An Alloy Which Melts in Hot Water

CADMIUM LEAD TIN BISMUTH

Bismuth lowers the melting point of alloys and so is used in alloys of tin and lead for making plugs for fire sprinklers. Alloys can be made with bismuth that will melt in hot water.

Wood's metal is a fusible alloy the melting point of which is about 160 degrees Fahrenheit. It will melt in contact with hot water.

In order to make Wood's metal you will need an accurate postal scale or a small chemical balance such as is used by amateur photographers. The alloy contains 8 parts of lead, 4 parts of tin, 15 parts of bismuth, and 3 parts of cadmium. It does not matter whether you weigh out the metals in pounds, ounces, or grams. Simply substitute whatever unit of weight you decide to use for the same number of parts. For example, use 8 ounces or 8 grams of lead, 4 ounces or 4 grams of tin, etc.

Melt the metals in a clean can or, better still, a clay crucible. Melt the lead first, add the cadmium, then the bismuth, and finally the tin, in the order named. Stir the mixture with a stick and then allow it to cool. If you have a mold for casting lead soldiers, you can cast some soldiers out of Wood's metal and mystify your friends. If a toy soldier cast from Wood's metal is dropped into a cup of very hot tea, the soldier will melt.

CHAPTER XI

PRACTICAL USES FOR YOUR CHEMICAL KNOWLEDGE

YOU can put your knowledge of chemistry to practical use in many ways.

For instance, you can use it in removing spots and stains from clothing, upholstery, rugs, and table linen. If you know what caused a spot or a stain, a knowledge of chemistry will often enable you to select a solvent which will dissolve the stain without causing any injury to the fabric.

Spots on clothing are most frequently caused by grease. The grease soon collects dust and dirt and shows plainly. Lubricating oil, vaseline, butter, cod-liver oil, lard, salad oil, gravy, and meat juices are some of the greasy substances which cause grease spots.

To remove a grease spot, you need a solvent which will *dissolve* the grease and *wash* out the dirt.

Professional cleaners use benzine, naphtha, gasoline, and other highly volatile liquids distilled from petroleum because these substances are effective solvents and are inexpensive. They are also DANGEROUS. They cause FIRES and EXPLOSIONS. Lighted cigarettes, matches, or stoves are not needed in order to ignite these liquids. They may be ignited by the spark in an electric-light switch or the static electricity generated in rubbing a spotted fabric. Avoid these dangerous liquids. Have nothing to do with them.

CARBON TETRACHLORIDE

Carbon tetrachloride is a safe cleaning fluid for grease spots. It is a light, colorless liquid used in such cleaning compounds as "Carbona." It is not only non-inflammable but is a component of chemical fire extinguishers such as "Pyrene." Carbon tetrachloride will dissolve oils, greases, asphalt, tar, rubber, and various gums, including chewing gum. It will also dissolve the sticky substances used on flypaper and adhesive tape.

AMYL ACETATE

Amyl acetate, commonly called banana oil because it has the odor of bananas, is another solvent useful for cleaning. It is a clear, colorless liquid used as a thinner and solvent for pyroxylin lacquers. It is also employed in varnishes and some paints. Amyl acetate readily dissolves resins, tannins, camphor, and celluloid. It may be used to remove spots of these substances and also nail polish and paint. It is inflammable and burns furiously, so be careful with it. Amyl acetate may be purchased at paint and hardware stores.

TURPENTINE

Turpentine is also a solvent. It will remove spots of fresh paint. Fresh paint, that is, paint which has not dried and hardened, is easier to remove than old paint. Old paint spots can sometimes be softened by soaking with turpentine. Paints vary so much in composition that it is impossible to select any one solvent as an ideal paint remover. It may be necessary to use several in order to secure results.

ALCOHOL

Alcohol, and by that is meant ethyl alcohol or the common beverage alcohol, is one of the best solvents. When ethyl alcohol is *denatured* or mixed with other substances to prevent its use as a beverage, it is called industrial alcohol. It will dissolve many varnishes, resins, enamels, soaps, drugs, essential oils, and shellac.

ACETONE

Acetone is a colorless, inflammable liquid with a mintlike odor. It is one of the most important of the industrial solvents and is used extensively in the manufacture of lacquers, celluloid, and smokeless powders. It has the power to absorb or dissolve many times its own volume of acetylene gas. Acetylene, it will be remembered, is made by the action of water on calcium carbide. When supplied with sufficient air, acetylene burns with a dazzling white light. It is often used for lighting in camps and places far removed from gas factories. The chief industrial use of acetylene is in the oxy-acetylene torch which can attain a temperature of 6,300 degrees Fahrenheit. The torch is employed for cutting and welding materials. Practically all the acetylene employed for welding is stored by dissolving in acetone. Acetone may be used as a cleaner to remove spots or stains caused by many varieties of lacquers, varnishes, enamels, and paints.

Solvents are not the only substances useful in removing spots and stains. Although a suitable solvent may be successful in dissolving and removing most of the objectionable substance which caused a spot, there is often a stain left which can be obliterated only by bleaching.

HYDROGEN PEROXIDE

Hydrogen peroxide is a bleach. Hydrogen and oxygen combine to form two compounds, one, our old friend water, containing 2 atoms of hydrogen and 1 atom of oxygen; and hydrogen peroxide, containing 2 atoms of hydrogen and 2 atoms of oxygen. By weight, water contains 1 part of hydrogen and 8 parts of oxygen; hydrogen peroxide contains 1 part of hydrogen and 16 parts of oxygen by weight.

Pure hydrogen peroxide is a clear, sirupy liquid which is so unstable that it explodes when exposed to heat or light. The ordinary variety to be found in the medicine cabinet, where it is kept for use as a gargle and an antiseptic for cuts and wounds, is a water solution containing about 3 per cent hydrogen peroxide.

Pure hydrogen peroxide holds its excess of oxygen very loosely, so loosely, in fact, that some of the oxygen is continually escaping. This also is true of the ordinary diluted variety used in the home. The solution slowly decomposes and forms water and free oxygen. If you pour some into the palm of your hand, you can see the bubbles of oxygen form. The bubbles are *nascent* oxygen. Nascent means that the atoms have been set free but have not yet combined to form molecules. In this state, the oxygen is highly active. That is why a solution of hydrogen peroxide is so useful as an antiseptic and a bleach. An ordinary household solution of hydrogen peroxide will bleach such articles as hair, silk, wool, feathers, and ivory.

The best results in bleaching a stain are secured if a few drops of household ammonia are added to the hydrogen peroxide. Before bleaching a stain, make a test on an inconspicuous portion of the fabric so as to determine whether or not the color of the material will fade.

SODIUM HYPOCHLORITE

The bleaching properties of chlorine have already been discussed.

The most convenient source of chlorine is the weak solution of sodium hypochlorite you can purchase at a grocery store under such names as Oxol and Clorox. This is a 5 per cent solution by weight and is too strong to use on fabrics without dilution. You can make a dilute solution by mixing from 1 to 3 tablespoons of a 5 per cent solution in 1 quart of cold water.

Remember that chlorine is used only for bleaching cotton goods and paper. Silk and wool are bleached by other means. If the stained fabric is starched, it will be necessary first to remove the starch by soaking in water.

The proper methods of removing such stains as those of coffee, tea, chocolate, eggs, fruit, and grass are quite well known. The procedure for more difficult stains caused by such substances as paint, tar, mercurochrome, iodine, etc., is not so familiar.

Below are directions for removing the more difficult types of stains.

THE TECHNIQUE OF STAIN REMOVAL

The important thing in stain removal is to treat the stain as soon as possible after it occurs; if possible, before it has had an opportunity to dry.

In several instances, the directions call for "sponging." To sponge a stained fabric, place it spot side *down* upon a clean, white blotter or several layers of white cotton cloth. Dip a small pad of clean, white cotton cloth into the solvent and rub the stained fabric lightly with it. Heavy rubbing will roughen some materials and is to be avoided. Move the spot to a new place upon the blotter or absorbent cloth from time to time. When the stain has loosened and started to disappear, turn the fabric over and sponge the spotted side lightly. In order to avoid forming a ring around the place where the spot was, complete the sponging with a cloth pad *barely moistened* with the solvent.

Bleaching is best accomplished by stretching the stained fabric over a bowl and dropping the bleaching fluid on the spot with a medicine dropper.

Use plenty of solvent. It is the solvent that cleans, not the rubbing.

Raise the cloth gradually at the finish of each stroke. This helps to avoid leaving a ring.

In general, it is hard to clean silk materials. Most silks are adulterated during manufacture by weighting or sizing with tin. This weighting makes it very difficult to avoid rings when cleaning. It is best to make a test on some inconspicuous portion of the material.

Adhesive Tape

When the surgical tape used to fasten bandages is removed, it usually leaves gummy black marks on the skin. These can be removed with a small soft pad which has been saturated with carbon tetrachloride or ether.

Bloodstains

Old bloodstains are hard to remove, so they should be treated as soon as possible after they occur. Sponge or wash immediately with cold water. Do not use hot water. Bloodstains are set and made more permanent by hot water. If any stain remains after the cold-water treatment, soak in a dilute ammonia solution and wash again.

Chewing Gum

The base of this sticky substance is chicle gum, made from the milky juice of the sapodilla tree. It may be softened with kerosene or turpentine and then sponged off with carbon tetrachloride.

Cod-Liver Oil

Sponge with banana oil or carbon tetrachloride first. Then use soapy water.

Fly-Paper

The sticky substance used on fly-paper is insoluble in water. It may be removed by sponging with carbon tetrachloride.

Grease

Practically all grease spots are soluble in carbon tetrachloride.

Ink

Inks vary considerably in their composition and there is no method of removing ink stains which can be applied to all inks.

Treat ink stains as soon as possible. If still wet, remove as much ink as possible with blotting paper. Then sponge or wash with hot, soapy water. The remaining stain can often be bleached with hydrogen peroxide or sodium hypochlorite solution. Remember that bleaches can not be used on colored fabrics without danger of fading the color. If a bleach is used, wash it off thoroughly.

Black drawing ink contains finely ground carbon and forms a stain which can not be bleached. A fresh spot of drawing ink should be blotted and washed with water. An old spot can be loosened with *spirits of camphor.*

Iodine

Tincture of iodine, used as an antiseptic for cuts and wounds, makes a yellowish-brown stain on unstarched materials and a blue stain on starched fabrics. Iodine stains may be removed by sponging with dilute ammonia or with wood alcohol. A solution of the "hypo" used by photographers will bleach iodine stains.

Lipstick

First sponge with carbon tetrachloride. Then treat with hydrogen peroxide or a weak solution of sodium hypochlorite.

Mercurochrome

Mercurochrome makes a stain which is difficult to remove. An old spot can seldom be entirely obliterated. Washing or sponging with cold water and bleaching with hydrogen peroxide or sodium hypochlorite solution produces the best results.

Nail Polish

Sponge with banana oil.

Oil

Use the methods recommended for grease.

Paint

Paints and enamels vary considerably in composition and consequently no one method of removal will prove ideal.

If it is a fresh spot, scrape off as much as possible with a dull knife. An old and hardened spot may be softened with turpentine, carbon tetrachloride, banana oil, or wood alcohol.

After scraping, soak thoroughly with one of the abovenamed solvents, using the one which seems to be most effective. A strong, hot solution of kitchen soap in water will then remove most of the loosened paint.

Traces of the colored pigment which remain can often be bleached with sodium hypochlorite solution.

Paraffin

Scrape off as much of the paraffin as possible with a dull knife and press between two layers of blotting paper as directed for candle wax.

Pencil

Do not use water on indelible pencil marks. Soak in denatured alcohol or sponge with same. Carbon tetrachloride is also effective. Bleach any stain remaining with sodium hypochlorite solution.

Rust

Rust is difficult to remove, especially from colored materials. The process of removing the rust usually fades the color. Rust is iron oxide. Lemon juice will remove faint rust spots. Oxalic acid does a better job but oxalic acid is very poisonous. A mixture of equal parts of powdered potassium binoxalate and powdered potassium bitartrate, moistened with enough water to form a thin paste and smeared on a rust spot, will bleach it. Do not use this mixture, however, on colored fabrics without first testing it on an inconspicuous place to ascertain whether or not it removes the color.

Shellac

See instructions for sealing-wax, under *Wax.*

Tar

Carbon tetrachloride will soften and dissolve tar.

Wax

Molten candle wax frequently drips on clothing and table linen. The wax itself is easily removed but the dye mixed with the wax used in colored candles is more difficult to obliterate.

Remove as much of the wax as possible by scraping with a dull knife. Then place the stained material between two pieces of clean blotting paper and press with a warm iron. The heat of the iron will melt the wax and it will be absorbed by the blotting paper. The remaining wax can usually be removed by sponging with carbon tetrachloride. If any stain from the coloring matter in the candle remains, try sponging it with wood alcohol. If dye stain still remains, it probably can be bleached with hydrogen peroxide. Hydrogen peroxide should not be used on colored fabrics because it will fade or bleach the material.

Sealing wax consists principally of shellac. It often contains resin also. Shellac is a resinous material which comes from the insect called *Tachardia lacca.*

To remove sealing wax, scrape off as much as possible with a dull knife and then sponge with alcohol.

USEFUL RECIPES AND FORMULAS

The amateur chemist can put chemistry to practical use in other ways besides removing spots and stains. He can mix many useful compounds and preparations.

For this work a graduate for measuring fluid ounces and a scale for weighing are needed. A small postal scale weighing 8 ounces and showing ¼-ounce divisions will be satisfactory.

Luminous Paint

This paint will absorb light and glow in the dark. It may be used to paint gate-posts, fuse-boxes, door frames, railings, electric light switches, or any object which must be located in the dark.

Grind 2½ ounces of barium sulfate in a mortar with 9 ounces of luminous calcium sulfide. Do not use home-made calcium sulfide, but purchase some from a chemical supply house. Be certain to obtain the luminous variety. When the two powders have been thoroughly mixed stir them into a one-quart can containing 12 fluid ounces of clear 4-hour varnish. Keep the can sealed when not in use. The surface to be painted must be thoroughly dry and clean. Water decomposes luminous sulfide compounds.

You can cut the quantity of each ingredient in half, or double or triple it, making as much of the paint as needed.

Cement for Fastening Porcelain and Metal Together

Add whiting to a strong solution of sodium silicate (water glass) and stir it well. Use enough whiting so that the mixture forms a thick cream. Use this as a cement to fasten porcelain to metal or vice versa. Allow the cement to harden for 48 hours before using the articles which have been fastened together.

Plumber's Cement

Mix 1 teaspoonful of the yellow monoxide of lead, called litharge, with enough glycerine to form a thick paste. Put this away for a day or two where it will be undisturbed. It will set and form a hard mass. A mixture of litharge and glycerine is called plumber's cement and is very useful in making repairs. It can be used to fasten porcelain handles to water faucets, seal drain-pipes to basins, etc.

It is necessary to make a fresh mixture of litharge and glycerine each time that you need some and to use it before it starts to harden.

To Kill Poison Ivy

Poison ivy can be killed without risk of touching the plants and becoming poisoned.

Dissolve 1½ pounds of sodium thiocyanate in 1 gallon of water. Put this solution in a sprayer and thoroughly saturate all the leaves and stems of the ivy plant above ground. The best time to apply the spray is in June.

Marine Putty

Ordinary glazier's putty is a mixture of whiting and linseed oil. When used to seal windows in their frames, it frequently becomes cracked and loosens. Ordinary putty is unsatisfactory for windows and skylights on a boat.

"White-lead" putty, made of a mixture of white lead and whiting, is harder than glazier's putty. It does not crack easily. It is used by boat builders. In fact it makes a better job of glazing any window or skylight.

To make white-lead putty, put some thick white lead ground in oil (such as painters use for mixing paint) on a clean board. Sprinkle a little whiting on the white lead. Then, using your fingers or a putty knife, knead and work the whiting into the white lead for several minutes. Add whiting from time to time until the mass has the same sticky consistency as ordinary putty. If too much whiting is used, the putty will not adhere properly when it is pressed into place and more white lead must be added.

Aquarium Cement

The glass sides of a rectangular aquarium or fish-tank are cemented into the metal framework. When a tank is not in use, the cement dries and sometimes cracks so that it leaks when filled with water. Then the tank must be taken apart, the old cement removed and fresh cement put in its place.

Some aquarists build their own tanks. It is not difficult to make cement suitable for a new tank or recementing an old one.

To 5 pounds of glazier's putty add ½ pound of dry litharge (yellow oxide of lead) and ½ pound of minium (red oxide of lead). Mix thoroughly and knead with boiled linseed oil until the mass has the consistency of thick putty. Add sufficient lampblack to give the cement a slate-gray color.

Sodium Silicate Cement

Paste two sheets of paper together, using sodium silicate solution as the cement. You will find that "water-glass" makes an exceptionally strong cement for paper and cardboard. It is used as such in making corrugated cardboard boxes.

A Safe Cleaning Fluid for Grease Spots

Ordinary benzine is an excellent solvent for grease but is dangerous to use as a cleaning fluid because both the liquid and its vapor are highly inflammable. A small quantity of benzine may cause a disastrous fire and explosion.

Benzine may be rendered safe and non-inflammable by mixing with carbon tetrachloride. Carbon tetrachloride is also a good grease solvent. The mixing should be done out of doors away from all fire. Use 3 parts of ordinary benzine and 1 part of carbon tetrachloride. Or, in other words, add 1 cupful of carbon tetrachloride to each 3 cupfuls of benzine. Put the mixture in a can or bottle with a tight stopper. It evaporates rapidly.

Ink Eradicator

Write or draw upon a piece of paper, using ordinary fountain pen ink. When the ink is dry, trace over the lines with a small camel's-hair brush which has been dipped in a solution consisting of two teaspoonfuls of water and one teaspoonful of sodium hypochlorite solution. This is sold at grocery stores under the name of Clorox.

You can use diluted sodium hypochlorite solution to remove ink blots and stains from cloth, drawings, books, papers, etc. The black ink will be bleached to a pale yellow.

Radiator Solder

Mix ½ pound of flaxseed meal with 1 ounce of aluminum powder. Stir it until each particle of meal is covered with aluminum.

When this mixture is added to the water in a leaky automobile radiator or cooling system, it swells and plugs up all leaks as the water circulates.

Anti-rot Compound for Wood

Dissolve ½ pound of sodium fluoride in 3 gallons of water. Soak the wood to be treated in this solution, or apply the mixture to the unpainted surfaces with a brush.

Silver Polish

Mix 1 teaspoonful of trisodium phosphate and 1 teaspoonful of sodium carbonate with 2 pounds of whiting. Add enough water to form a thin cream. Apply this with a soft rag to the metal to be cleaned and rub lightly. Let the compound remain on the metal until it dries and then polish with a soft cloth.

This is an excellent polishing compound for chromium plating.

Brass Polish

Thoroughly mix ¼ pound of tartaric acid with 2 pounds of fine tripoli powder. Wet a cloth, dip it in the powder and use it to polish brass, nickel, etc.

Polishing Cloths

Dissolve ½ ounce of oxalic acid in 2 quarts of water. Add ½ pound of whiting and stir thoroughly. Dip some soft cloths such as canton flannel in this mixture and when they are thoroughly saturated, squeeze them out and hang up to dry.

When dry, the cloths may be used to polish metal and glass.

Brass Polishing Paste

Mix ½ pound of whiting with 2 ounces of turpentine and 1 ounce of alcohol. Add enough household ammonia to form a paste. A rag smeared with this paste will polish brass.

Transfer Solution

Pictures from the funny papers and cartoons can be transferred to a sheet of white paper with this solution.

Mix 2 teaspoonfuls of tincture of green soap and 2 teaspoonfuls of turpentine with 1 quart of water.

Dampen the newsprint picture with this solution. This will soften the ink. Lay the picture face down on a sheet of white paper and rub the back with the bowl of a silver-plated tablespoon. When the picture is peeled off, a facsimile will remain on the white sheet. Several copies of the same picture can be made in this manner.

Red Ink

Dissolve ¼ ounce of the dye called EOSINE in 1 quart of water. Add ¼ ounce of powdered gum arabic and a few drops of phenol (carbolic acid).

Stamp Pad Ink

A strong solution of dye in water can be used as stamp pad ink but quickly dries. By adding glycerine, the ink can be prevented from evaporating too rapidly. The following dyes are suitable for making ink: nigrosine (black), soluble blue, light green, magenta red, and acid violet.

Starch Paste

Ordinary starch paste is prepared by mixing starch with a small amount of cold water and then adding boiling water.

The paste may be considerably improved by adding glycerine and ammonium hydroxide. Ammonium hydroxide increases the strength of the paste and glycerine renders it flexible.

Leather Cleaner

Dissolve 3 ounces of powdered Castile soap in 5 pints of boiling water. When the soap has dissolved add 3 ounces of ammonium hydroxide, 7 ounces of glycerine and 4 ounces of ethylene dichloride.

CHAPTER XII

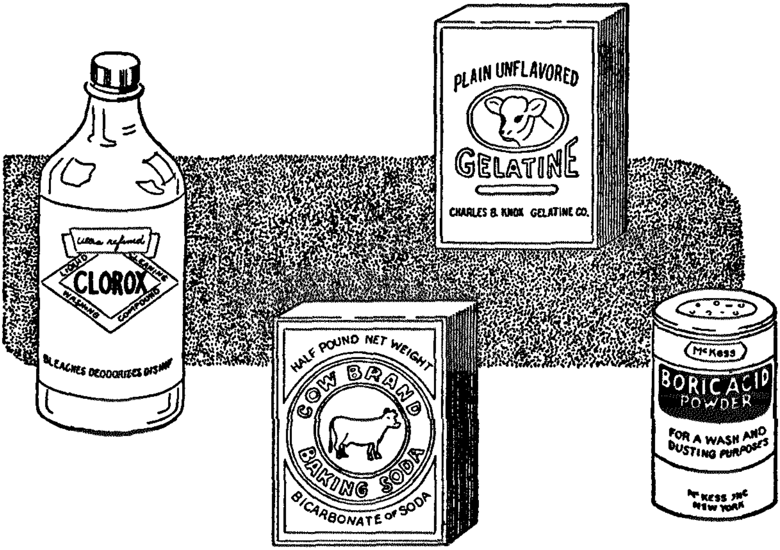
THE CHEMICALS YOU WILL NEED

A COMPLETE list of all the chemicals required to perform all the experiments described in this book is given below. You do not need all these to perform most of the experiments. Some of them are required for one or two experiments only. With others you can perform a great many experiments.

A few of the chemicals are everyday substances in common use, to be found in the kitchen or medicine chest. Others can be bought at a drug-store.

Many you will have to purchase from a dealer in laboratory supplies. There are such firms in every large city. You can find who they are and where they are located by consulting the classified telephone directory. Several concerns which deal in chemicals for experimenters advertise in the popular mechanical and scientific magazines. They will fill orders for small quantities of chemicals by mail. It is possible to buy 5 cents' worth of many compounds from some of these firms. Look in the classified columns for their advertisements.

Keep your chemicals in tightly corked bottles and not in paper, wood, or metal containers. A great many compounds decompose under the action of light. Your chemicals should not be kept where sunlight will fall upon them. All bottles should be labeled to show their contents.



Some of the chemicals used to perform the experiments in this book are common substances in the kitchen and medicine chest.

Chemical Experimenting Is Not Dangerous

Some persons think it is. They are wrong. If you spill a can of paint, it will do damage. If you drink it, paint will poison you. Paint will catch fire and burn. But it would be silly to be afraid of a small can of paint or fear to use the paint.

It is silly to be afraid of ordinary chemicals or to consider that all forms of chemical experiments are dangerous.

All the experiments described in this book and all the chemicals required are safe if used and handled intelligently. The medicine cabinet of nearly every home contains substances more dangerous than any of those recommended for your laboratory.

A sharp knife or a chisel is dangerous in the hands of a child. They may be so if *improperly handled* by an adult. But that is no reason why an intelligent boy should not use a knife or a chisel.

When chemicals are used in small quantities and used in the proper manner, they do no harm. If you eat a cake of soap, it will make you very ill. If you wash your hands with soap, it will do no harm. Eating it is obviously an improper use of soap. A bottle of tincture of iodine can be found in the medicine cabinet of most homes. The bottle is marked "poison." If you put it in your mouth, it will burn and poison you. If you spill it on clothing, it will stain. But if you use it properly, daub a little bit on cuts and scratches, it is beneficial.

The rules of common sense and intelligence in handling chemicals and experiments are simple. Here they are.

1. Unless you are 14 years of age or older, do not perform experiments using acids, sodium hydroxide, or potassium chlorate. If you wish to perform experiments with these substances, ask an older boy or an adult to assist and guide you.
2. If you spill chemicals, wipe them up immediately.
3. Do not make up any experiments of your invention. Do not mix any chemicals or unknown materials unless you know it is safe to do so.
4. Do not taste chemicals or put them in your mouth.
5. Do not get chemicals on your hands, clothes, skin, or furniture. Always wash your hands after experimenting.

6. Keep your chemicals in a safe place where small children can not get hold of them.

7. Your best safeguard is to follow instructions explicitly and literally.

Here is a list of the chemicals used to perform the experiments in this book. The chemical formula for each substance is given and in many instances the approximate cost and source of supply. The chemically pure (abbreviated C. P.) grade is the most expensive. The technical grade (abbreviated Tech.) costs less than the C. P. variety and is satisfactory for any of the experiments described here.

ALUM ($K_2Al_2(SO_4)_3$). There are a great many alums but the term is commonly used to mean a double sulfate of potassium and aluminum. Powdered alum may be purchased at any drugstore.

ALUMINUM POWDER (Al). The fine flakes of aluminum used in paint and for sale at paint and hardware stores are satisfactory for use in the experiments described in this book. The cost of aluminum powder at a laboratory-supply house is about 10 cents for 1 ounce.

ALUMINUM SULFATE ($Al_2(SO_4)_3$). White crystals, soluble in water. Crude aluminum sulfate is called "alum-cake" and is used in the purification of sewage. The cost of ¼ pound is about 10 cents at a laboratory-supply house.

AMMONIUM CARBONATE ($(NH_4)_2CO_3$). This is an unstable compound which decomposes at ordinary temperatures, forming ammonia and ammonium bicarbonate. It is sometimes called "lump" or "rock" ammonia. May be purchased at some grocery stores. Many drug-stores also carry it in stock. Available from any dealer in laboratory supplies. The cost of chemically pure ammonium carbonate is about 20 cents for ¼ pound.

Ammonium Chloride (NH_4Cl). Obtainable from a laboratory-supply house. A commercial grade suitable for making batteries and performing the experiments described in this book may be purchased at hardware and paint stores under the name of sal ammoniac. The cost of the technical grade is about 10 cents for ¼ pound.

Ammonium Hydroxide (NH_4OH). May be purchased at a grocery store in the form of household ammonia. Pure ammonium hydroxide is clear. The cloudiness in the household variety is due to impurities. This will not impair its usefulness for most home-laboratory experiments. Ammonium hydroxide is ammonia gas united with water. Unless it is kept tightly stoppered, the gas will escape and the solution will become weakened. The bottles in which household ammonia is sold are usually provided with rubber corks or stoppers. Rubber stoppers seal a bottle more tightly than vegetable corks. The ammonium hydroxide which the chemist uses is more concentrated than household ammonia and is kept in glass-stoppered bottles.

Ammonium Nitrate (NH_4NO_3). A white crystalline salt used as an ingredient of fireworks and explosives. Obtainable from a dealer in laboratory supplies at a cost of about 20 cents for ¼ pound.

Ammonium Sulfide ($(NH_4)_2S_x$) is a liquid used by chemists in analytical work. It is inexpensive and may be purchased from a laboratory-supply house.

Barium Chloride ($BaCl_2$). Obtainable from a laboratory-supply house at a cost of approximately 15 cents for ¼ pound.

Bismuth (Bi). A grayish-white metal which also has a reddish tinge. The cost of 1 ounce is about 30 cents at a laboratory-supply house.

Borax ($Na_2B_4O_7$). May be purchased at a grocery store. Large quantities of borax are used in the manufacture of enamels, glazes, and optical glass. Borax occurs as a natural deposit in the dried-up lakes in some parts of California, India, and Tibet.

Boric Acid (H_3BO_3) is also known as boracic acid and may be purchased at any drug-store.

Cadmium (Cd). In color, this metal resembles tin. Obtainable at a laboratory-supply house at a cost of about 35 cents for 1 ounce.

Calcium Carbonate ($CaCO_3$). The chemist's name for chalk. Powdered chalk may be purchased at any drug-store. The cost of ¼ pound should not be over 10 cents.

Calcium Chloride ($CaCl_2$). This substance is produced in large quantities as a by-product in several industrial operations. Fused or dehydrated calcium chloride is a porous mass which is used in chemical laboratories for drying gases and liquids. Since it is a highly deliquescent substance and is inexpensive, calcium chloride is often sprinkled on concrete work to attract water from the atmosphere and assist in the setting process. May frequently be purchased at a drug-store or builder's-supply house. Laboratory-supply houses charge 10 cents for ¼ pound.

Calcium Hydroxide ($Ca(OH)_2$). May be prepared by mixing calcium oxide or quicklime with water. This is called slaking the lime.

Slaked lime (calcium hydroxide) for making plaster is sold at lumber yards and dealers in masons' materials. A plasterer will probably be glad to give you a small amount. Seed stores also carry slaked lime in stock for "liming" lawns and gardens. The technical grade of calcium hydroxide sold by laboratory-supply houses costs about 15 cents for ¼ pound.

Calcium Oxide (CaO). Is sometimes obtainable at hardware stores in small cans under the name of *quicklime* for making whitewash. Calcium oxide must be kept in a sealed container because, when exposed to the air, it gradually absorbs moisture and carbon dioxide, forming calcium carbonate. *Do not handle* *calcium oxide with your fingers.* It will react with the moisture in your skin and burn. The technical variety costs 10 cents for ¼ pound.

CALCIUM SULFIDE (CaS). The luminous variety of this substance is an impure grade. It costs about 30 cents per ounce at a laboratory-supply house.

CAMPHOR. A vegetable gum. May be purchased at a drug-store.

CARBON DISULFIDE (CS_2). Pure carbon disulfide is colorless and its odor is not unpleasant. Commercial carbon disulfide is a yellow liquid. Its smell is usually disagreeable. It is a good solvent for rubber, waxes, fats, gums, resins, bromine, camphor, sulfur, iodine, and phosphorus. May be purchased at a drug-store. The cost is about 25 cents for 4 ounces.

CARBON TETRACHLORIDE (CCl_4). A heavy, colorless liquid which is a safe fluid for dry-cleaning. It does not burn and its vapor will smother fires. It may be purchased at hardware and automobile-supply stores. Carbon tetrachloride is very volatile and must be kept tightly sealed or it will evaporate. Costs about 15 cents for 4 ounces.

CHARCOAL (C). Blocks of charcoal for use with a blowpipe cost 10 cents each at a laboratory-supply house. Charcoal can be purchased in large bags at some hardware and grocery stores.

COBALTOUS CHLORIDE ($CoCl_2$). A ruby-red crystalline substance which costs about 50 cents for ¼ pound at laboratory-supply houses.

CONGO RED. An aniline dye used for making test paper. Obtainable from a dealer in laboratory chemicals at a cost of about 30 cents for 1 ounce.

COPPER SULFATE ($CuSO_4$). May be obtained from any laboratory-supply house and occasionally from drug-stores. An impure variety known as blue vitriol and bluestone is sold by hardware and seed stores for killing weeds. If you purchase this impure or crude variety, pick out the dark azure-blue crystals from the light-blue powder with which they are mixed. Preserve these crystals in a stoppered bottle. The cost in small quantities is about 10 cents for ¼ pound.

CUPRIC NITRATE ($Cu(NO_3)_2$). Obtainable from a laboratory-supply house at a cost of approximately 35 cents for ¼ pound.

CUPRIC OXIDE (CuO). Cupric oxide is a black substance, usually in the form of a powder. It is formed by heating copper in an atmosphere of oxygen. It can also be made by igniting copper carbonate, nitrate, or hydroxide. Copper oxide may be purchased from a dealer in laboratory chemicals at a cost of approximately 25 cents for ¼ pound of the technical grade.

ETHYL ALCOHOL (C_2H_5OH). The young experimenter will find it difficult to purchase this variety of alcohol because of government restrictions controlling its sale. Perhaps a doctor or chemist friend who knows that you are not going to drink it will help you to obtain a small amount for experiments.

FERROUS SULFATE ($FeSO_4$). Green crystals called green vitriol or copperas, soluble in water. Ferrous sulfate is efflorescent, that is, loses its water of crystallization unless kept in a tightly corked bottle. The technical grade, obtainable at a laboratory-supply house, costs about 15 cents for ¼ pound.

GLYCERINE ($C_3H_5(OH)_3$). Known to the chemist as *glycerol.* It is obtainable at any drug-store. The cost is about 30 cents for 4 ounces.

HYDROCHLORIC ACID (HCl). Should be kept in a glass-stoppered bottle and treated with respect. May be purchased at drug-stores and from dealers in laboratory chemicals. The usual cost of 4 ounces is about 35 cents.

HYDROGEN PEROXIDE (H_2O_2). Is obtainable at any drug-store. You can purchase a small bottle at the dime store for 10 cents. This variety put up for home use is a water solution containing about 3 per cent hydrogen peroxide. Pure hydrogen peroxide is a clear, syrupy liquid about 1½ times as heavy as water. It is so unstable that it explodes when exposed to heat or light.

When hydrogen peroxide is called for in any of the experiments in this book, it is the ordinary household variety that is intended. This is perfectly safe to handle. However, if spilled, it may bleach the color out of any fabrics with which it comes in contact. Hydrogen peroxide must be kept in a tightly stoppered *brown* bottle in a cool, dark place or it will lose its oxygen and the solution will revert to water.

IODINE (I). Iodine crystals may be purchased at a drug-store or laboratory-supply house. The cost of the pharmaceutical grade is approximately 40 cents per ounce. Keep iodine crystals in a tightly stoppered bottle.

IODINE, TINCTURE OF. Obtainable at any drug-store. You can also purchase a small bottle for 10 cents at the dime store.

IRON OXIDE (Fe_2O_3). May be purchased from a dealer in laboratory chemical supplies. The cost of ¼ pound is usually about 10 cents.

LEAD ACETATE ($Pb(C_2H_3O_2)_2$). This is a white crystalline compound but is usually sold in powdered form. It is also called sugar of lead. Unlike most basic salts, lead acetate is soluble in water. It is poisonous. May be purchased at drug-stores and chemical-supply houses. The cost is about 15 cents for ¼ pound.

LEAD ACETATE PAPER. May be purchased from a laboratory-supply house or made by staining white absorbent paper with a solution of lead acetate.

LEAD CHROMATE ($PbCrO_4$). Lead chromate is used as a yellow pigment in paints.

LEAD OXIDE. See *Lead peroxide, Litharge,* and *Red oxide of lead.*

LEAD PEROXIDE (PbO_2). A brown powder which is a strong oxidizing agent. If hydrogen sulfide is allowed to come into contact with a few grains of lead peroxide, the gas is oxidized and bursts into flame. The cost of 1 ounce at a laboratory-supply house is 25 cents.

Lime-water ($Ca(OH)_2$). A solution of calcium hydroxide in water which can be purchased at a drug-store or made in the home laboratory by shaking up some calcium hydroxide with water, and filtering.

Litharge (PbO). The common name for the yellow powder which is yellow oxide of lead. Obtainable at some drug-stores and from dealers in laboratory chemicals. The cost of ¼ pound is about 20 cents.

Litmus Cubes. Used in preparing litmus-paper. The cubes are dissolved in hot water and the solution used to stain white unsized paper. Obtainable at a laboratory-supply house. One ounce of litmus cubes costs about 20 cents.

Litmus Paper. Prepared litmus-paper is inexpensive and usually is obtainable at a drug-store.

Logwood. Also called Campeachy wood. A natural dyestuff once of great commercial importance but replaced in recent years by artificial coloring matters. Sold by laboratory-supply houses in the form of dark-red or brown chips which produce a deep-red solution when soaked in hot water. One ounce costs about 20 cents.

Magnesium, Powdered (Mg). Used in making landing flares, photographers' flashes, and fireworks. Price is approximately 30 cents per ounce at laboratory-supply houses.

Magnesium Sulfate ($MgSO_4$). Is inexpensive and may be purchased at any drug-store under the name of Epsom salts.

Magnesium Ribbon (Mg). A silver-white lustrous metallic ribbon produced from magnesium chloride by electrolysis. Burns with an intense white light. A 15-foot length costs 10 cents at a laboratory-supply house.

Manganese Dioxide (MnO_2). A black powder, insoluble in water. Used commercially on a large scale in the manufacture of dry cells. Obtainable at laboratory-supply houses at a cost of about 10 cents for ¼ pound.

MANGANESE SULFATE ($MnSO_4$). Rose-colored crystals, soluble in water, which may be purchased from a laboratory-supply firm at approximately 15 cents for ¼ pound.

MERCURIC NITRATE ($Hg(NO_3)_2$). A yellowish, crystalline substance obtainable from a laboratory-supply house at a cost of about 50 cents for 1 ounce.

MERCURIC OXIDE (HgO). A red powder. All mercury compounds are somewhat expensive. Mercuric oxide is sold by dealers in laboratory supplies at approximately 80 cents for ¼ pound.

METHYL ORANGE. A dye used as a reagent. Obtainable from a dealer in laboratory chemicals at a cost of about 50 cents for 1 ounce.

NITRIC ACID (HNO_3). Should be kept in a glass-stoppered bottle and treated with respect. Never spatter or spill nitric acid or allow it to come into contact with skin, clothing, furniture, etc. May be obtained at a drug-store or from a dealer in laboratory chemicals. The cost of 4 ounces is usually about 35 cents.

NICKEL AMMONIUM SULFATE ($(NH_4)_2SO_4.NiSo_4$). Bluish-green crystals used in nickel plating and laboratory analysis. Obtainable from a dealer in platers' supplies or from a laboratory-supply house. Costs 15 cents for ¼ pound.

OXONE (Na_2O_2). A special brand of sodium peroxide manufactured for producing oxygen. See *Sodium peroxide.*

PHENOLPHTHALEIN. Phenolphthalein is a powder which dissolves in alcohol and is used as a test for alkalis. It turns red in the presence of alkalis. Obtainable at drug-stores or laboratory-supply houses. The cost of 1 ounce is approximately 20 cents. One ounce is sufficient for the average amateur laboratory.

PHOSPHORUS, RED (P). Red phosphorus is a dull-red powder consisting of small crystals. It is obtained by heating yellow phosphorus in a vessel from which air is excluded. It is not poisonous and, unlike yellow phosphorus, does not have to be kept under water to avoid spontaneous combustion. The cost of red phosphorus from a dealer in laboratory chemicals is approximately 20 cents for 1 ounce or 35 cents for ¼ pound.

POTASSIUM CHLORATE ($KClO_3$). Powdered potassium chlorate may be obtained at a drug-store or a chemical-supply house. Two or 3 ounces will be sufficient to perform all the experiments in which it is called for in this book. The cost is about 30 cents for ¼ pound. Obtain powdered potassium chlorate. *Do not attempt to powder it yourself. It is dangerous to do so.* Potassium chlorate must be clean or it is explosive. If it gets dirt in it, throw it away. Clean potassium chlorate is perfectly safe.

POTASSIUM DICHROMATE ($Na_2Cr_2O_7$). Orange-red crystals soluble in water. Can be purchased at a drug-store sometimes. The technical grade, obtainable at a laboratory-supply house, costs about 25 cents for ¼ pound.

POTASSIUM FERROCYANIDE ($K_4Fe(CN)_6$). A yellowish crystalline compound also known as yellow prussiate of potash. Costs approximately 30 cents for ¼ pound at a laboratory-supply house.

POTASSIUM IODIDE (KI). When potassium iodide crystallizes, it forms large opaque cubes. Silver iodide is used in medicine and photography. In the laboratory potassium iodide is used whenever an iodide is required. May be purchased at a drug-store or from a chemical-supply house. The cost is about 25 cents for 1 ounce.

POTASSIUM MANGANATE ($KMnO_4$). Obtainable only from a laboratory-supply house. Do not confuse this with potassium permanganate. It is a different substance.

POTASSIUM NITRATE (KNO_3). Also called saltpeter. A white crystalline substance used to "corn" beef, in medicine, gunpowder, fireworks, and as a fertilizer. Obtainable at a drug-store. Laboratory-supply firms charge about 15 cents for ¼ pound.

Potassium Permanganate ($KMnO_4$). Purple crystals with a greenish luster which dissolve in water, forming a solution with an intense purple color. *Potassium permanganate crystals should not be handled with the fingers.* Obtainable at a drugstore or a laboratory-supply house at a cost of approximately 25 cents for ¼ pound.

Red Oxide of Lead (Pb_3O_4). A red powder, also called minium, obtainable from a dealer in laboratory chemicals at a cost of 15 cents for ¼ pound.

Silver Nitrate ($AgNO_3$). A white crystalline substance, also known as lunar caustic. *Do not handle silver nitrate crystals with the fingers or get any of the solution on your skin.* All silver salts are somewhat expensive. An ounce of silver nitrate will cost approximately 75 cents. May be obtained at a drugstore.

Sodium Bisulfate ($NaHSO_4$). These colorless or white crystals are frequently used by young experimenters to replace sulfuric acid. Sodium bisulfate is safe for the novice to handle, whereas sulfuric acid is not. One-quarter of a pound costs 10 to 15 cents at a laboratory-supply house.

Sodium Bicarbonate ($NaHCO_3$). May be purchased in very pure form at any grocery store under the name "baking soda." You can buy a good-sized box for 10 cents. Keep it in a wide-mouthed bottle or a covered glass jar.

Sodium Carbonate (Na_2CO_3). A dry white powder, sometimes called soda ash. Another variety, in the form of large transparent crystals, is sold in grocery stores as "washing soda." Washing soda may be used in any of the experiments in this book calling for sodium carbonate. On exposure to the air, crystals of washing soda gradually lose their water of crystallization and crumble into a white powder (soda ash). A technical grade of sodium carbonate is sold by laboratory-supply houses at 10 cents per ¼ pound.

SODIUM HYDROXIDE ($NaOH$). Pure sodium hydroxide may be obtained at a chemical-supply house at a cost of about 25 cents per ¼ pound. The lye sold in grocery and hardware stores is impure sodium hydroxide. The impurities (usually about 6 per cent) will not impair its usefulness for most home-laboratory experiments. Sodium hydroxide must be kept in a tightly closed can or jar. Otherwise it will not only absorb moisture from the air, but will be converted gradually into sodium carbonate by atmospheric carbon dioxide. *Treat sodium hydroxide with respect always. Do not handle it with your fingers or spill it.* Sodium hydroxide will dissolve skin and destroy fabrics. If you spill any or get it on your fingers, use vinegar or some other mild acid to counteract it.

SODIUM HYPOCHLORITE ($NaOCl$). May be purchased at grocery stores in the form of a 5 per cent solution known under the trade-name of Clorox.

SODIUM PEROXIDE (Na_2O_2). A yellowish-tinged substance (commercial grade) which must be kept in a tightly stoppered bottle. It decomposes in moist air. *Always use a spoon to handle sodium peroxide. Do not touch it with the fingers.* The cost at a laboratory-supply house is 30 cents for ¼ pound.

SODIUM SILICATE SOLUTION (Na_2SiO_3). Also known as water glass. May be purchased at some drug-stores. A thick, colorless or yellowish liquid which sells in small quantities at 15 cents for 4 ounces.

SODIUM SULFATE (Na_2SO_4). Known to druggists as Glauber's salt. One quarter pound sells for about 10 cents.

SODIUM TETRABORATE ($Na_2B_4O_7$). See *Borax.*

SODIUM THIOSULFATE ($Na_2S_2O_3$). Sold by dealers in photographic supplies under the name of hypo. It is inexpensive. Buy a 1-pound package and put the contents into a glass jar which can be kept sealed. Otherwise your hypo will absorb moisture from the air.

STRONTIUM NITRATE ($Sr(NO_3)_2$). Strontium compounds are used in fireworks to make "red fire." Laboratory-supply firms charge about 15 cents for ¼ pound.

SULFUR (S). May be purchased at drug-stores and chemical-supply houses in the form of a flaky yellow powder known as flowers of sulfur. Some hardware and seed stores also carry flowers of sulfur. The cost is about 10 cents per ¼ pound.

SULFURIC ACID (H_2SO_4). Should be kept in a glass-stoppered bottle and *treated with respect.* Never pour water into acid. The cost of dilute acid is about 15 cents for 4 ounces. The same quantity of concentrated acid costs about 35 cents. May be purchased at drug-stores and from dealers in laboratory chemicals.

TIN (Sn). Tin is a silver-white crystalline metal used as a coating on mild steel to make tin-plate and tin cans. Obtainable at a laboratory-supply house at a cost of about 20 cents for 1 ounce.

ZINC POWDER (Zn). A bluish-gray powder obtainable at a laboratory-supply house at a cost of about 15 cents for ¼ pound.

CHEMICAL FORMULAS

In writing the names of chemicals, the chemist finds it convenient to use abbreviations for the names of elements and compounds. Then it can be seen at a glance just what the elements are and how they combine. The abbreviations the chemist uses are called "symbols." Each element has its own symbol. The various combinations of the symbols of the elements that compose a compound are called "formulas." Each symbol indicates a single atom of an element unless it has a small number following. The number represents the number of atoms of that element in the molecule.

Below is a list of the most common laboratory chemicals and their formulas. When you label your bottles, include the chemical formula as well as the name. Become familiar with the formulas of the chemicals which you use.

Acetaldehyde	CH_3CHO
Acetamide	CH_3CONH_2
Acetanilid	$C_6H_5NHCOCH_3$
Acetone	CH_3COCH_3
Acetyl chloride	CH_3COCl
Acid acetic	CH_3COOH
Acid arsenic	As_2O_5
Acid arsenous	As_2O_3
Acid benzoic	C_6H_5COOH
Acid boric	B_2O_3
Acid butyric	$CH_3(CH_2)_2COOH$
Acid carbolic	C_6H_5OH
Acid chromic	CrO_3
Acid cinnamic	$C_6H_5(CH)_2COOH$
Acid citric	$H_3C_6H_5O_7+H_2O$
Acid formic	$HCOOH$
Acid gallic	$C_6H_2(OH)_3COOH+H_2O$
Acid hydriodic	HI
Acid hydrobromic	HBr
Acid hydrochloric	HCl
Acid hydrofluoric	HF
Acid iodic	HIO_3
Acid lactic	$CH_3CHOHCOOH$
Acid nitric	HNO_3
Acid oleic	$C_{17}H_{33}COOH$
Acid osmic	OsO_4
Acid oxalic	$H_2C_2O_4+2H_2O$
Acid palmitic	$CH_3(CH_2)_{14}COOH$
Acid perchloric	$HClO_4$
Acid phosphoric	H_3PO_4
Acid phosphorous	H_3PO_3
Acid phthalic	$C_6H_4(CO)_2O$

Acid picric	$C_6H_2(NO_2)_3OH$
Acid pyrogallic	$C_6H_3(OH)_3$
Acid salicylic	$C_6H_4OHCOOH$
Acid stearic	$CH_3(CH_2)_{16}COOH$
Acid succinic	$(CH_2COOH)_2$
Acid sulfuric	H_2SO_4
Acid sulfurous	H_2SO_3
Acid tannic	$C_{14}H_{10}O_9$
Acid tartaric	$(CHOHCOOH)_2$
Acid trichloracetic	CCl_3COOH
Alcohol amyl	$CH_3CHCH_2CH_2OH$
Alcohol benzyl	$C_6H_5CH_2OH$
Alcohol butyl	$CH_3(CH_2)_2CH_2OH$
Alcohol ethyl	C_2H_5OH
Alcohol methyl	CH_3OH
Alcohol propyl iso	$(CH_3)_2CHOH$
Aluminum	Al
Aluminum acetate	$Al(CH_3COO)_2OH$
Aluminum ammonium sulfate	$Al_2(NH_4)_2(SO_4)_4+24H_2O$
Aluminum carbide	Al_4C_3
Aluminum chloride	$AlCl_3+6H_2O$
Aluminum hydroxide	$Al(OH)_3$
Aluminum nitrate	$Al(NO_3)_3+9H_2O$
Aluminum oxide	Al_2O_3
Aluminum potassium sulfate	$Al_2K_2(SO_4)_4+24H_2O$
Aluminum sodium sulfate	$Al_2Na_2(SO_4)_4.24H_2O$
Aluminum sulfate	$Al_2(SO_4)_3+18H_2O$
Aluminum sulfide	Al_2S_3
Ammonium acetate	NH_4OOCCH_3
Ammonium bicarbonate	NH_4HCO_3
Ammonium bichromate	$(NH_4)_2Cr_2O_7$
Ammonium bisulfate	NH_4HSO_4
Ammonium bisulfite	NH_4HSO_3
Ammonium bromide	NH_4Br
Ammonium chloride	NH_4Cl
Ammonium chromate	$(NH_4)_2CrO_4$
Ammonium citrate	$(NH_4)_2HC_6H_5O_7$

Ammonium fluoride	NH_4F
Ammonium hydroxide	NH_4OH
Ammonium iodide	NH_4I
Ammonium nitrate	NH_4NO_3
Ammonium nitrite	NH_4NO_2+Aq
Ammonium oxalate	$(NH_4)_2C_2O_4+H_2O$
Ammonium persulfate	$(NH_4)_2S_2O_8$
Ammonium sulfate	$(NH_4)_2SO_4$
Ammonium sulfide	$(NH_4)_2S_x$
Ammonium tartrate	$(NH_4)_2C_4H_4O_6$
Ammonium thiocyanate	NH_4SCN
Amyl acetate	$C_5H_{11}CO_2CH_3$
Aniline	$C_6H_5NH_2$
Antimony	Sb
Antimony trioxide	Sb_2O_3
Antimony sulfate	$Sb_2(SO_4)_3$
Antimony pentasulfide	Sb_2S_5
Antimony trisulfide	Sb_2S_3
Arsenic	As
Arsenic chloride	$AsCl_3$
Arsenic sulfide (Red)	As_2S_2
Arsenic sulfide (Yellow)	As_2S_3

Barium	Ba
Barium acetate	$Ba(C_2H_3O_2)_2+H_2O$
Barium bromide	$BaBr_2+2H_2O$
Barium carbonate	$BaCO_3$
Barium chlorate	$Ba(ClO_3)_2+H_2O$
Barium chloride	$BaCl_2+2H_2O$
Barium hydroxide	$Ba(OH)_2$
Barium nitrate	$Ba(NO_3)_2$
Barium peroxide	BaO_2
Barium sulfate	$BaSO_4$
Barium sulfide	BaS
Benzaldehyde	C_6H_5CHO
Benzene	C_6H_6
Benzyl chloride	$C_6H_5CH_2Cl$

Bismuth	Bi
Bismuth chloride	$BiCl_3$
Bismuth nitrate	$Bi(NO_3)_3+5H_2O$
Bismuth subnitrate	$BiONO_3+H_2O$
Bismuth trioxide	Bi_2O_3
Bromine	Br_2

Cadmium	Cd
Cadmium acetate	$Cd(C_2H_3O_2)_2+2H_2O$
Cadmium carbonate	$CdCO_3$
Cadmium chloride	$2CdCl_2+5H_2O$
Cadmium iodide	CdI_2
Cadmium nitrate	$Cd(NO_3)_2+4H_2O$
Cadmium potassium iodide	$K_2CdI_4+2H_2O$
Cadmium sulfate	$3CdSO_4+8H_2O$
Calcium	Ca
Calcium acetate	$Ca(C_2H_3O_2)_2+H_2O$
Calcium bromide	$CaBr_2$
Calcium carbide	CaC_2
Calcium carbonate	$CaCO_3$
Calcium chloride	$CaCl_2$
Calcium cyanamide	$CaCN_2$
Calcium fluoride	CaF_2
Calcium hydroxide	$Ca(OH)_2$
Calcium hypochlorite	$CaOCl_2$
Calcium nitrate	$Ca(NO_3)_2+4H_2O$
Calcium oxide	CaO
Calcium phosphate monobasic	$Ca(H_2PO_4)_2+H_2O$
Calcium phosphate dibasic	$CaHPO_4$
Calcium phosphate tribasic	$Ca_3(PO4)_2$
Calcium phosphide	Ca_3P_2
Calcium sulfate	$CaSO_4+2H_2O$
Calcium sulfide	CaS
Calcium sulfite	$CaSO_3+2H_2O$
Carbon disulfide	CS_2
Carbon tetrachloride	CCl_4

Chloroform	$CHCl_3$
Chromium	Cr
Chromium acetate	$Cr(C_2H_3O_2)_3+Aq$
Chromium ammonium sulfate	$Cr(NH_4)_2(SO_4)_4+24H_2O$
Chromium chloride	$CrCl_3+6H_2O$
Chromium nitrate	$Cr(NO_3)_3+9H_2O$
Chromium oxide	Cr_2O_3
Chromium potassium sulfate	$K_2Cr_2(SO_4)_4+24H_2O$
Chromium sulfate	$Cr_2(SO_4)_3+Aq$
Cobalt	Co
Cobalt acetate	$Co(C_2H_3O_2)_2+4H_2O$
Cobalt chloride	$CoCl_2+6H_2O$
Cobalt nitrate	$Co(NO_3)_2+6H_2O$
Cobalt oxide	Co_3O_4
Cobalt sulfate	$CoSO_4+7H_2O$
Copper	Cu
Cupric acetate	$Cu(C_2H_3O_2)_2+H_2O$
Cupric ammonium chloride	$Cu(NH_4)_2Cl_4+2H_2O$
Cupric bromide	$CuBr_2$
Cupric carbonate	$CuCO_3Cu(OH)_2$
Cupric chloride	$CuCl_2+2H_2O$
Cupric nitrate	$Cu(NO_3)_2+3H_2O$
Cupric oxide	CuO
Cupric sulfate	$CuSO_4+5H_2O$
Cupric sulfide	CuS
Cuprous bromide	Cu_2Br_2
Cuprous chloride	Cu_2Cl_2
Cuprous cyanide	$Cu_2(CN)_2$
Cuprous oxide	Cu_2O
Cuprous sulfide	Cu_2S

Dextrose	$C_6H_{12}O_6+H_2O$
Dichlorobenzene	$C_6H_4Cl_2$

Ether	$(C_2H_5)_2O$
Ethyl acetate	$CH_3COOC_2H_5$

Ethyl benzoate	$C_6H_5COOC_2H_5$
Ethyl bromide	C_2H_5Br
Ethyl iodide	C_2H_5I
Ethylene glycol	$(CH_2OH)_2$
Ferric ammonium sulfate	$FeNH_4(SO_4)_2+12H_2O$
Ferric chloride	$FeCl_3+6H_2O$
Ferric citrate	$FeC_6H_5O_7+5H_2O$
Ferric-ferrous oxide	Fe_3O_4
Ferric nitrate	$Fe(NO_3)_3+9H_2O$
Ferric oxalate	$Fe_2(C_2O_4)_3+6H_2O$
Ferric oxide	Fe_2O_3
Ferric sulfate	$Fe_2(SO_4)_3+Aq$
Ferrous ammonium sulfate	$Fe(NH_4)_2(SO_4)_2+6H_2O$
Ferrous chloride	$FeCl_2+4H_2O$
Ferrous oxalate	$FeC_2O_4+2H_2O$
Ferrous sulfate	$FeSO_4+7H_2O$
Ferrous sulfide	FeS
Fluorescein	$C_{20}H_{12}O_5$
Formaldehyde	$HCHO$
Furfural	C_4H_3OCHO
Glycerine	$C_3H_5(OH)_3$
Gold	Au
Gold chloride	$HAuCl_4+3H_2O$
Hydrogen peroxide	H_2O_2
Hydroquinone	$C_6H_4(OH)_2$
Iodine	I_2
Iodoform	CHI_3
Iron	Fe
Lactose	$C_{12}H_{12}O_{11}+H_2O$
Lead	Pb

Lead acetate	$Pb(C_2H_3O_2)_2+3H_2O$
Lead carbonate	$Pb(OH)_2.2PbCO_3$
Lead chloride	$PbCl_2$
Lead chromate	$PbCrO_4$
Lead nitrate	$Pb(NO_3)_2$
Lead oxide (Yellow)	PbO
Lead oxide (Red)	Pb_3O_4
Lead peroxide	PbO_2
Lead sulfate	$PbSO_4$
Lead sulfide	PbS
Levulose	$C_6H_{12}O_6$
Lithium	Li
Lithium carbonate	Li_2CO_3
Lithium chloride	$LiCl+H_2O$
Lithium nitrate	$LiNO_3$
Lithium sulfate	$Li_2SO_4+H_2O$
Magnesium	Mg
Magnesium acetate	$Mg(C_2H_3O_2)_2+4H_2O$
Magnesium chloride	$MgCl_2+6H_2O$
Magnesium nitrate	$Mg(NO_3)_2+6H_2O$
Magnesium oxide	MgO
Magnesium sulfate	$MgSO_4$
Manganese	Mn
Manganese chloride	$MnCl_2+4H_2O$
Manganese dioxide	MnO_2
Manganese nitrate	$Mn(NO_3)_2$
Manganese sulfate	$MnSO_4+4H_2O$
Mercury	Hg
Mercuric acetate	$Hg(C_2H_3O_2)_2$
Mercuric bromide	$HgBr_2$
Mercuric chloride	$HgCl_2$
Mercuric cyanide	$Hg(CN)_2$
Mercuric iodide	HgI_2
Mercuric nitrate	$Hg(NO_3)_2$
Mercuric oxide (Red)	HgO
Mercuric oxide (Yellow)	HgO

Mercuric sulfate	$HgSO_4$
Mercuric sulfide	HgS
Mercurous chloride	$HgCl$
Mercurous iodide	HgI
Mercurous nitrate	$HgNO_3+H_2O$
Mercurous sulfate	Hg_2SO_4
Methyl acetate	CH_3COOCH_3
Methyl iodide	CH_3I
Molybdenum	Mo

Nickel	Ni
Nickel ammonium sulfate	$Ni(NH_4)_2(SO_4)_2+6H_2O$
Nickel chloride	$NiCl_2+6H_2O$
Nickel hydroxide	$Ni(OH)_2$
Nickel nitrate	$Ni(NO_3)_2+6H_2O$
Nickel oxide	Ni_2O_3
Nickel sulfate	$NiSO_4+6H_2O$

Paraformaldehyde	$(HCHO)_3$
Paraldehyde	$(CH_3CHO)_3$
Phenolphthalein	$(C_6H_4OH)_2CC_6H_4COO$
Phosphorus	P
Phosphorus oxychloride	$POCl_3$
Phosphorus pentachloride	PCl_5
Phosphorus trichloride	PCl_3
Platinum	Pt
Platinum chloride	$PtCl_6+6H_2O$
Potassium	K
Potassium acetate	$KC_2H_3O_2$
Potassium arsenate	KH_2AsO_4
Potassium arsenite	$KAsO_2$
Potassium bicarbonate	$KHCO_3$
Potassium bichromate	$K_2Cr_2O_7$
Potassium bisulfate	$KHSO_4$
Potassium bisulfite	$K_2S_2O_5$
Potassium bitartrate	$KHC_4H_4O_6$
Potassium bromate	$KBrO_2$

Potassium bromide	KBr
Potassium carbonate	$K_2CO_3+H_2O$
Potassium chlorate	$KClO_3$
Potassium chloride	KCl
Potassium chromate	K_2CrO_4
Potassium citrate	$K_3C_6H_5O_7+H_2O$
Potassium cyanate	$KCNO$
Potassium cyanide	KCN
Potassium ferricyanide	$K_3Fe(CN)_6$
Potassium ferrocyanide	$K_4Fe(CN)_6+3H_2O$
Potassium fluoride	$KF+2H_2O$
Potassium hydroxide	KOH
Potassium iodate	KIO_3
Potassium iodide	KI
Potassium nitrate	KNO_3
Potassium nitrite	KNO_2
Potassium oxalate	$K_2C_2O_4+H_2O$
Potassium perchlorate	$KClO_4$
Potassium permanganate	$KMnO_4$
Potassium persulfate	$K_2S_2O_8$
Potassium phosphate monobasic	KH_2PO_4
Potassium phosphate dibasic	K_2HPO_4
Potassium phosphate tribasic	K_3PO_4+Aq
Potassium sulfate	K_2SO_4
Potassium sulfide	K_2S
Potassium sulfite	$K_2SO_3+2H_2O$
Potassium tartrate	$K_2C_4H_4O_6+\frac{1}{2}H_2O$
Potassium thiocyanate	$KSCN$
Pyridine	$CH(CHCH)_2N$
Quinhydrone	$C_6H_4O_2C_6H_4(OH)_2$
Resorcinol	$C_6H_4(OH)_2$
Silver	Ag
Silver acetate	$AgC_2H_3O_2$

Silver bromide	$AgBr$
Silver carbonate	Ag_2CO_3
Silver chloride	$AgCl$
Silver cyanide	$AgCN$
Silver iodide	AgI
Silver nitrate	$AgNO_3$
Silver oxide	Ag_2O
Silver sulfate	Ag_2SO_4
Sodium	Na
Sodium acetate	$NaC_2H_3O_2$
Sodium ammonium phosphate	$NaNH_4HPO_4+4H_2O$
Sodium arsenate	$Na_2HAsO_4+7H_2O$
Sodium arsenite	$NaAsO_2$
Sodium benzoate	$NaC_7H_5O_2$
Sodium bicarbonate	$NaHCO_3$
Sodium bichromate	$Na_2Cr_2O_7+2H_2O$
Sodium bisulfate	$NaHSO_4+H_2O$
Sodium bisulfite	$NaHSO_3$
Sodium bitartrate	$NaHC_4H_4O_6+H_2O$
Sodium borate	$Na_2B_4O_7+10H_2O$
Sodium bromate	$NaBrO_3$
Sodium bromide	$NaBr$
Sodium carbonate	Na_2CO_3
Sodium chlorate	$NaClO_3$
Sodium chloride	$NaCl$
Sodium chromate	$Na_2CrO_4+4H_2O$
Sodium citrate	$Na_3C_6H_5O_7+5\frac{1}{2}H_2O$
Sodium cyanide	$NaCN$
Sodium ferrocyanide	$Na_4Fe(CN)_6+12H_2O$
Sodium fluoride	NaF
Sodium formate	$NaOOCH$
Sodium hydroxide	$NaOH$
Sodium iodate	$NaIO_3$
Sodium iodide	NaI
Sodium molybdate	$Na_2MoO_4+2H_2O$
Sodium nitrate	$NaNO_3$
Sodium nitrite	$NaNO_2$
Sodium oxalate	$Na_2C_2O_4$

Sodium perborate	$NaBO_3+4H_2O$
Sodium peroxide	Na_2O_2
Sodium phosphate monobasic	$NaH_2PO_4+H_2O$
Sodium phosphate dibasic	$Na_2HPO_4+12H_2O$
Sodium phosphate tribasic	$Na_3PO_4+12H_2O$
Sodium potassium tartrate	$NaKC_4H_4O_6+4H_2O$
Sodium pyrophosphate	$Na_4P_2O_7+10H_2O$
Sodium salicylate	$NaC_7H_5O_3$
Sodium silicate	$Na_2SiO_3+5H_2O$
Sodium stannate	$Na_2SnO_3+3H_2O$
Sodium sulfate	$Na_2SO_4+10H_2O$
Sodium sulfide	NA_2S+9H_2O
Sodium sulfite	NA_2SO_3
Sodium tartrate	$NA_2C_4H_4O_6+2H_2O$
Sodium thiocyanate	$NaSCN$
Sodium thiosulfate	$Na_2S_2O_3+5H_2O$
Sodium tungstate	$Na_2WO_4+2H_2O$
Stannic chloride	$SnCl_4+5H_2O$
Stannic oxide	SnO_2
Stannous chloride	$SnCl_2+2H_2O$
Stannous oxide	SnO
Strontium acetate	$Sr(C_2H_3O_2)_2+\frac{1}{2}H_2O$
Strontium carbonate	$SrCO_3$
Strontium chloride	$SrCl_2+6H_2O$
Strontium nitrate	$Sr(NO_3)_2$
Strontium sulfate	$SrSO_4$
Sucrose	$C_{12}H_{22}O_{11}$
Sulfur	S
Thallium sulfate	Tl_2SO_4
Thorium	Th
Tin	Sn
Titanium	Ti
Titanium dioxide	TiO_2
Titanium tetrachloride	$TiCl_4$
Toluene	$C_6H_5CH_3$
Tungsten	W

Uranium	U
Uranium acetate	$UO_2(C_2H_3O_2)_2+2H_2O$
Uranium nitrate	$UO_2(NO_3)_2+6H_2O$
Uranium oxide	Na_2UO_4
Urea	$(NH)_2CO$

Xylene	$C_6H_4(CH_3)_2$

Zinc	Zn
Zinc acetate	$Zn(C_2H_3O_2)_2+2H_2O$
Zinc chloride	$ZnCl_2$
Zinc nitrate	$Zn(NO_3)_2+6H_2O$
Zinc oxide	ZnO
Zinc sulfate	$ZnSO_4+7H_2O$
Zinc sulfide	Zns

INDEX OF THE CHEMICALS REQUIRED TO PERFORM THE EXPERIMENTS DESCRIBED IN THIS BOOK

Look up in this index any chemicals which you have on hand. If you find them here, the numbers will refer you to the pages where you will find experiments utilizing these substances. The names of the chemicals required are printed in small capitals at the top of each experiment.

GENERAL INDEX