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THE GOLDEN BOOK OF CHEMISTRY EXPERIMENTS

HOW TO SET UP A HOME LABORATORY—OVER 200 SIMPLE EXPERIMENTS

THE GOLDEN BOOK OF

Chemistry Experiments

How to Set Up a Home Laboratory— Over 200 Simple Experiments

BY ROBERT BRENT

ILLUSTRATED BY HARRY LAZARUS

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Words Used by Chemists

Acid: a hydrogen-containing compound that releases hydrogen ions in solution.

Alloy: a material made up by combining two or more metals.

Analysis: breaking down a compound into two or more substances.

Anhydrous: free from water.

Atom: the smallest unit of an element that can enter into the making of a chemical compound.

Atomic weight: the weight of an atom compared with the weight of an oxygen atom set at 16.

Base: a compound containing the hydroxide group (OH).

Catalyst: a substance that helps in a chemical reaction without itself being changed.

Chemical change: a change of a substance into another substance having different properties.

Chemistry: a branch of science dealing with the compositions of substances and the changes that can be made in them.

Combustion: burning; a chemical change that produces heat and light.

Compound: a substance consisting of two or more different kinds of atoms in definite proportions by weight.

Crystal: a solid in which atoms or molecules are arranged in a definite pattern.

Density: the weight of a liquid or a solid in grams per cm^3 or milliliter.

Distillate: a liquid that has been turned into vapor and again cooled into a liquid.

Distillation: the process of producing a distillate.

Ductile: capable of being drawn out into a wire.

Electrolysis: breaking down a substance by passing an electric current through it.

Electrolyte: a substance that, when in a solution or when melted, will conduct an electric current.

Element: a substance that contains only one kind of atoms.

Equation: a complete description of a chemical reaction by the use of symbols, formulas, and signs.

Evaporation: the changing of a substance into vapor; also the process of removing water by heating.

Filtrate: a liquid obtained by filtration.

Filtration: the process of straining a liquid from a solid through porous material, usually filter paper.

Formula: a group of symbols and numbers giving the composition of a compound.

Hydrate: a compound containing loosely bound water of hydration (water of crystallization) that can be driven off by heating.

Hydroxide: a compound that contains the hydroxyl (OH) radical.

Ion: an electrically charged atom or group of atoms (radical).

Malleable: capable of being hammered or rolled into a thin sheet.

Matter: anything that takes up space and has weight.

Metal: an element that is a good conductor of electricity, has luster, and whose oxide forms a base with water.

Metalloid: an element that has properties of both metals and nonmetals.

Mixture: a mingling of substances not combined chemically.

Molecular weight: the sum of the atomic weights of the atoms that make up a molecule of a compound.

Molecule: the smallest unit of a compound that can exist in the free state.

Neutralization: the reaction of an acid and a base to give a salt and water.

Nonmetal: an element that is a poor conductor of electricity, does not have luster, and whose oxide forms an acid when combined with water.

Organic chemistry: the chemistry of the carbon compounds.

Oxidation: the process by which a substance combines with oxygen.

Precipitate: an insoluble solid formed in a solution by chemical reaction.

Radical: a group of atoms that behave chemically as a single atom.

Reaction: a chemical change.

Reduction: removal of oxygen; the opposite of oxidation.

Salt: compound (other than water) formed by the reaction of an acid and a base.

Saturated solution: a solution that contains the maximum amount of solute under the conditions.

Solubility: the number of grams of a solute needed to make a saturated solution in 100 grams of solvent.

Solute: the substance dissolved in a solvent.

Solution: a non-settling mixture of a solute in a solvent.

Solvent: a liquid in which a solute is dissolved.

Sublimation: a process by which a solid is turned into vapor and again cooled into a solid without passing through a liquid stage.

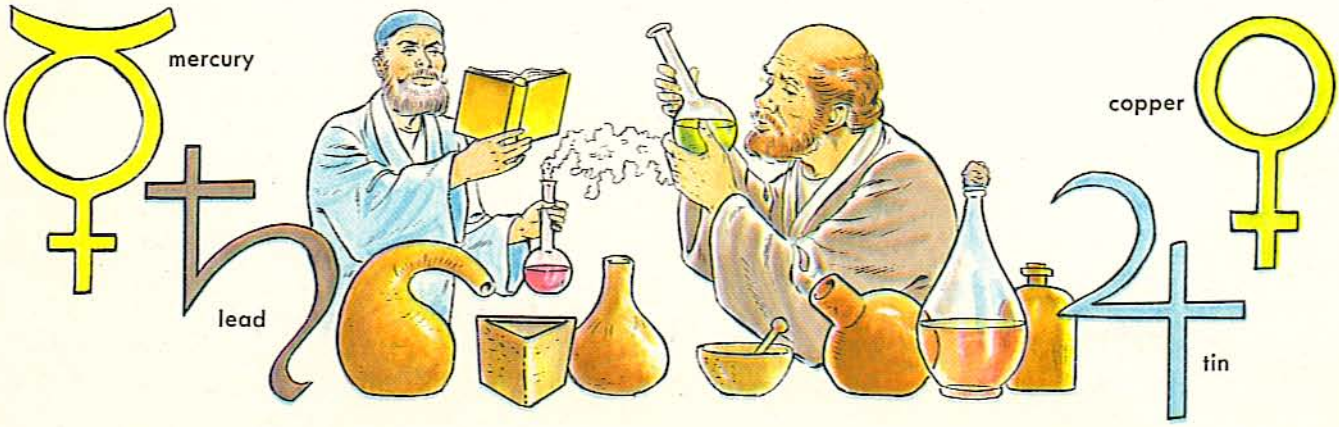
Subscript: a small numeral indicating the number of atoms of a certain element in the formula of a compound.

Substance: any specific kind of matter whether element, compound, or mixture.

Symbol: a letter or two letters representing one atom of an element.

Synthesis: the making up of a compound from simpler compounds or from elements; the opposite of analysis.

Valence: the number of hydrogen atoms which one atom of an element can displace or with which it can unite.



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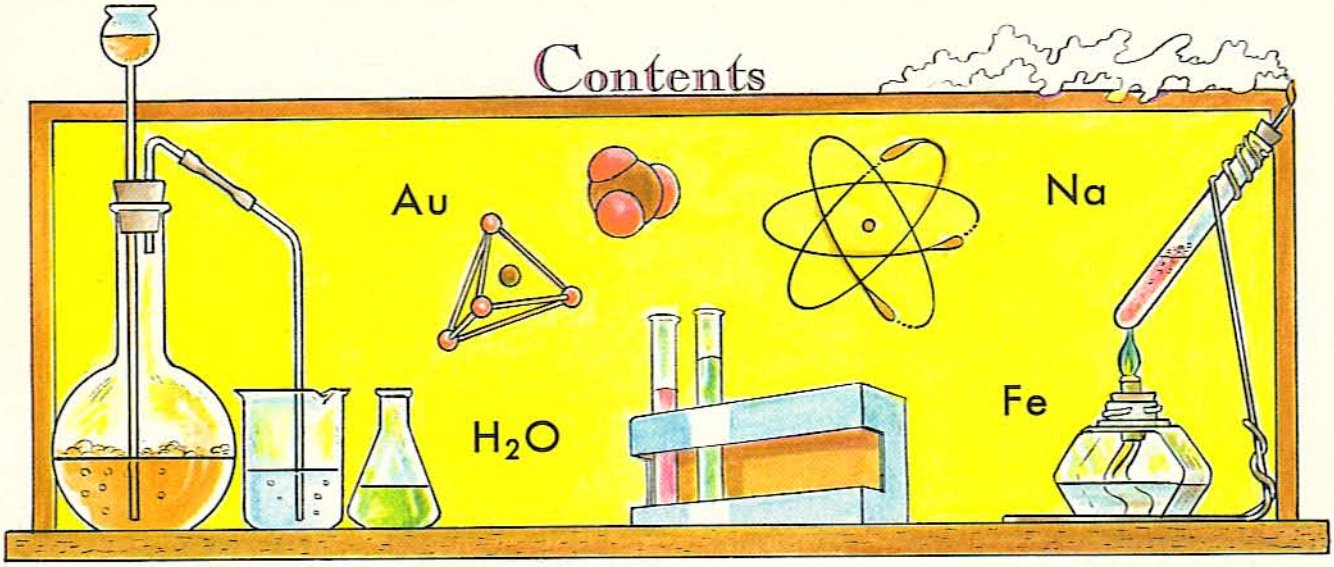

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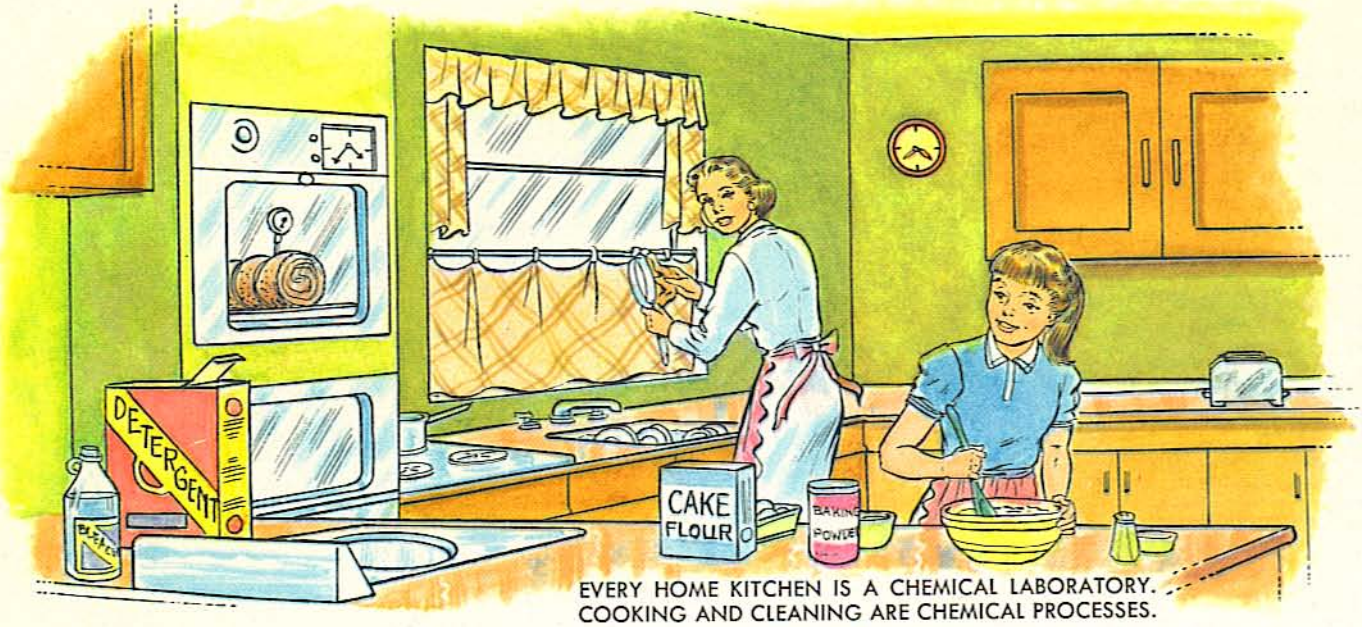


EVERY HOME KITCHEN IS A CHEMICAL LABORATORY. COOKING AND CLEANING ARE CHEMICAL PROCESSES.

The Importance of Chemistry

THERE IS HARDLY a boy or a girl alive who is not keenly interested in finding out about things. And that's exactly what chemistry is: FINDING OUT ABOUT THINGS — finding out what things are made of and what changes they undergo.

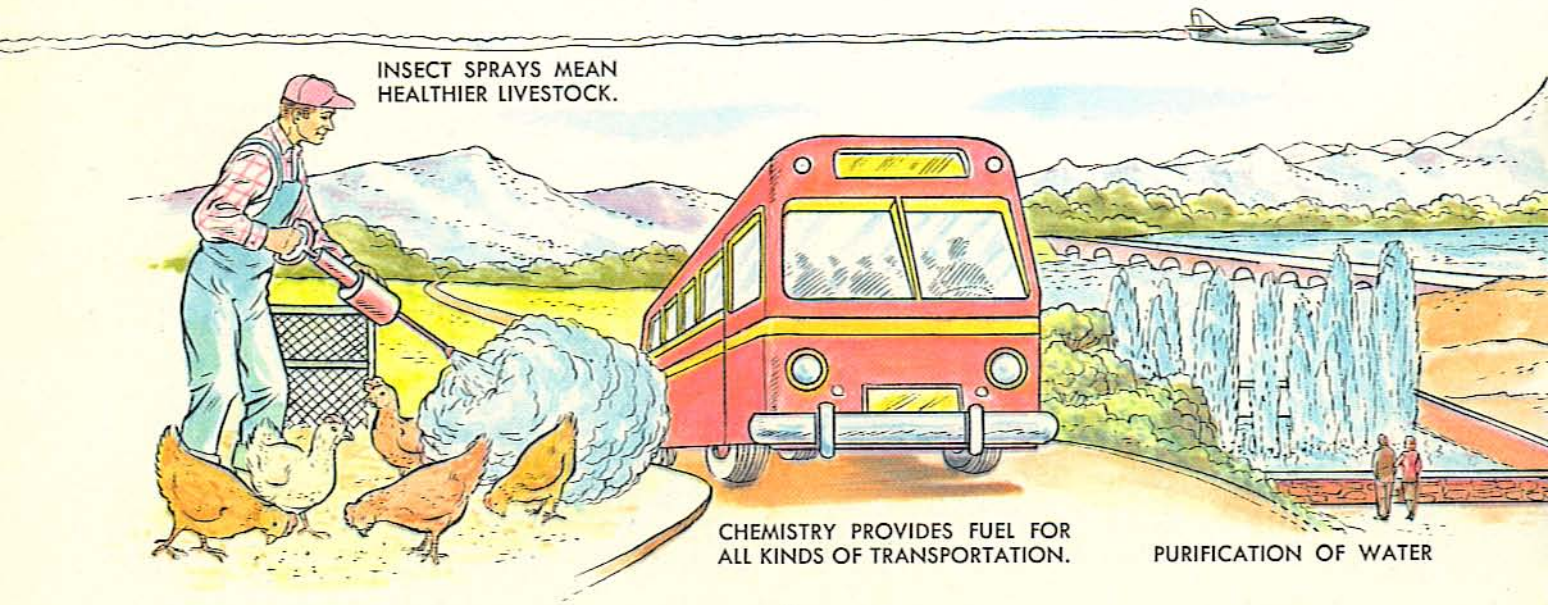
What things? Any thing! Every thing!

Take a look around you. All the things you see — and lots of things you can't see — have to do with the science of chemistry.

Let's start with yourself. The air you breathe is a mixture of chemical substances and the process of breathing is a chemical reaction. The foods you eat are all chemical products and the ways in which your body turns them into muscles and bones and nerves and brain cells are some of the greatest of all chemical mysteries.

The clothes you wear, the books you read, the medicine you take, the house in which you live — all these are products of chemistry. So is the family car — the metal in it, the rubber on which it rolls, the gas that moves it.

Nature itself is a tremendous chemical laboratory. Everything in nature is forever passing through chemical changes. Here on earth, plants and animals



INSECT SPRAYS MEAN HEALTHIER LIVESTOCK.

CHEMISTRY PROVIDES FUEL FOR ALL KINDS OF TRANSPORTATION.

PURIFICATION OF WATER

grow, die, and decay; rocks crack and crumble under the influence of air and water. In the universe, new stars are formed, others fade. The sun that gives us heat and light and energy is a flaming furnace of chemical processes that will eventually burn itself out, billions of years from now.

Chemistry is one of the most important of all sciences for human welfare.

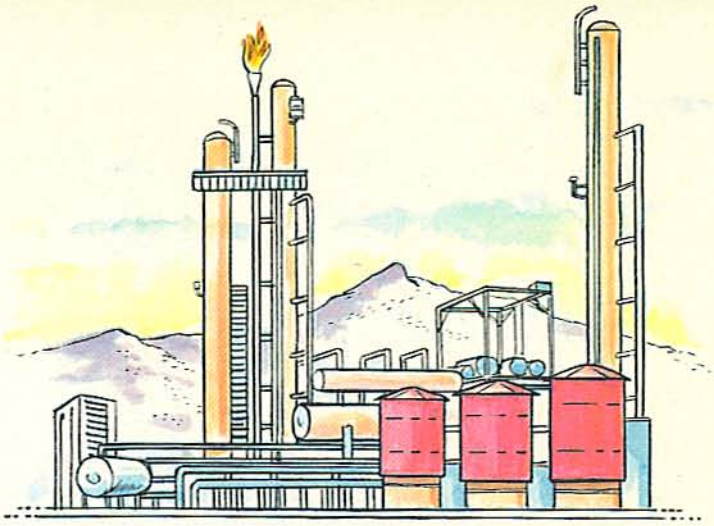
Chemistry means the difference between poverty and starvation and the abundant life. The proper use of chemistry makes it possible for farmers to feed the world's ever-increasing population, for engineers to develop new means of transportation and communication that will bring the peoples of the world closer together, for doctors to cure the diseases of mankind, for manufacturers to produce the thousands of items that are necessary for better and richer living.

And this is only the beginning.

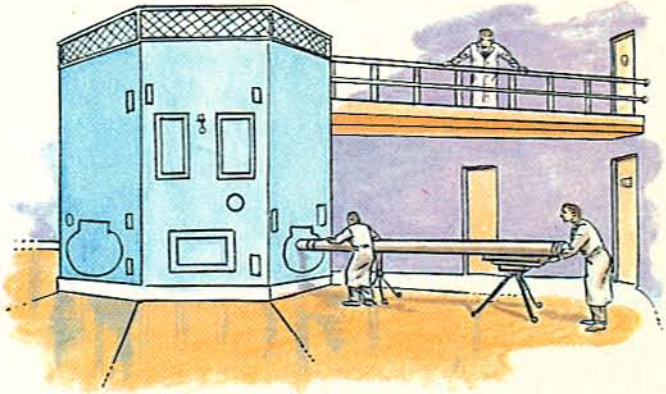
Within recent years, scientists have succeeded in penetrating into the innermost secrets of chemical substances and have begun to make use of the tremendous force that lies hidden in them. This atomic power opens up amazing possibilities for the future.

You will live in a world in which chemistry will become ever more important. To understand that world it is necessary to understand the truths and laws on which modern chemistry is based and to learn how chemists of the past unraveled them.

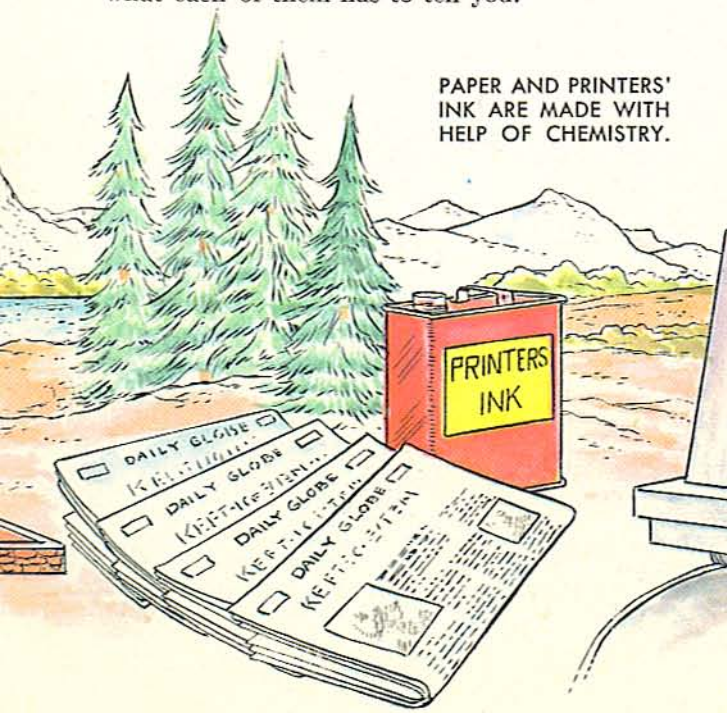
This book will help you get this insight — not alone by your reading it, but also by your conscientiously doing the experiments described and learning what each of them has to tell you.



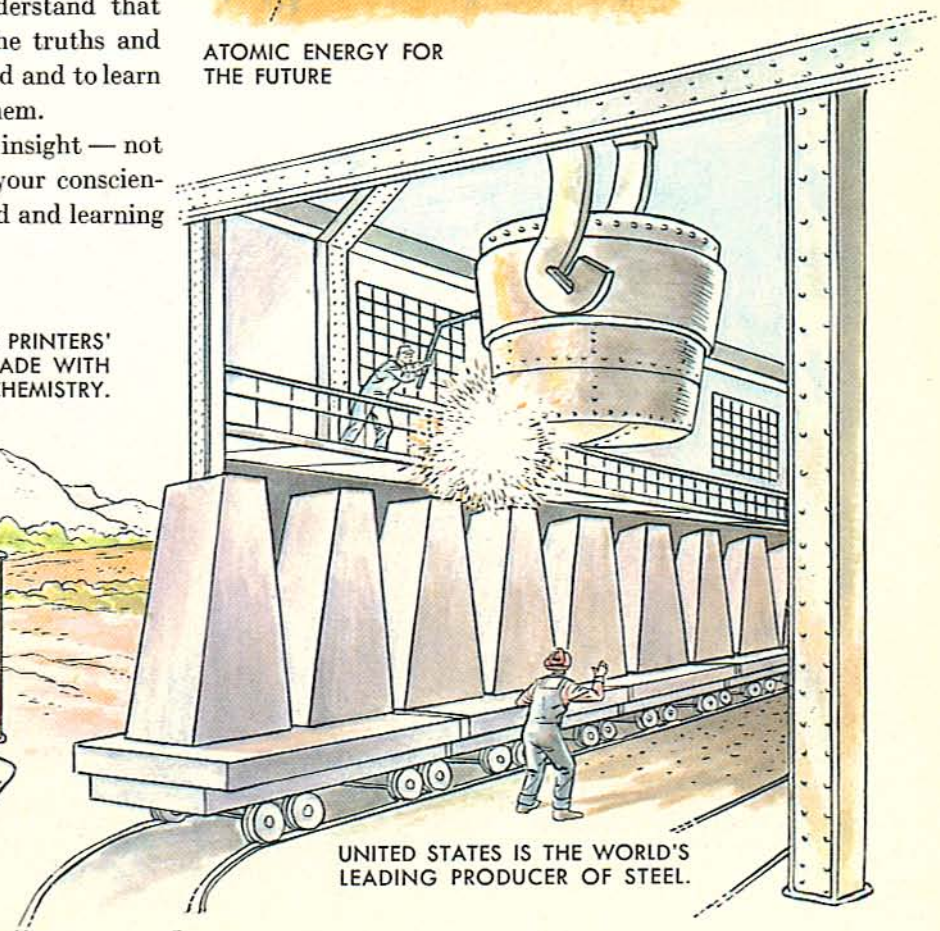
OIL IS THE BASIS FOR COUNTLESS CHEMICAL PRODUCTS.



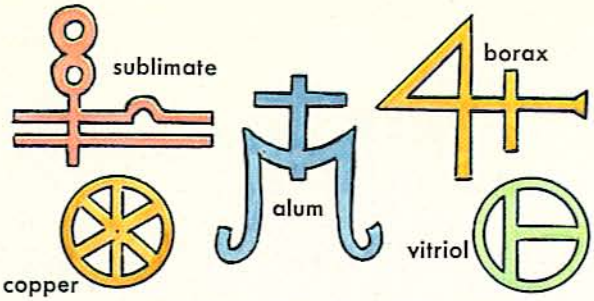
ATOMIC ENERGY FOR THE FUTURE



PAPER AND PRINTERS' INK ARE MADE WITH HELP OF CHEMISTRY.



UNITED STATES IS THE WORLD'S LEADING PRODUCER OF STEEL.



SYMBOLS USED BY ALCHEMISTS

Chemists of the Past

MANY THOUSAND years ago, an early ancestor of yours pushed a stick into the hot lava flowing from an erupting volcano. The stick burst into fire. He held it up as a torch. It gave off light and heat and finally turned into ashes.

This ancient man might be considered the world's first chemist. He had actually taken a substance called wood and had, by a chemical process called combustion or burning, turned it into something else.

The discovery of the use of fire was the first great step leading toward modern chemistry. Fire made it possible to turn raw foodstuffs into edible meals, to bake shaped clay into pottery, to make glass, to drive metals out of their ores.

For thousands of years people were chiefly interested in the results of what they did — they didn't care about what happened or why it happened. It was only about 2,500 years ago that philosophers began to wonder about what things were made of and what happened when a thing changed into something else.

Around 400 B. C., in Greece, a thinker by the name of Empedocles came up with an idea that seemed to make sense. He explained that everything in the world was made from just four things which he called "elements": fire, water, air, and earth. Think of that burning stick mentioned above. It gave off fire — so, obviously, the stick had to contain fire. It sizzled — which meant there was water in it. It smoked — and smoke would be some kind of air. It left ashes — and ashes are earth, as certainly everyone should know.

Everyone — except another Greek, Democritus, born around the time when Empedocles died. He had a different notion — that all matter was made up of tiny particles which he called *atomos* — something that cannot be cut further.

But Democritus didn't get very far with his idea. The greatest Greek philosopher of the day, Aristotle, held out for the four elements. And because of his great reputation this false idea governed the thinking of scientists for two thousand years — because no one dared suggest that he knew better than the great Aristotle!

BRONZE-AGE MAN WAS ONE OF THE EARLIEST CHEMISTS.

DEMOCRITUS INSISTED THAT MATTER CONSISTS OF ATOMS.

PARACELSUS TOLD HIS PUPILS TO USE EXPERIMENTS.

BOYLE INVESTIGATED GASES AND BROKE OLD TRADITIONS.

In the meantime, scientists of Arabia began work in a subject they called alchemy — from Arabic *al*, the, and *kimia*, pouring together. They mixed things and boiled and distilled and extracted in the hope, some day, of finding a way of making GOLD! They discovered a great number of things not previously known, developed many sound laboratory methods, and gave the science of chemistry its name — but they never created the slightest speck of gold. Neither did a great number of European alchemists.

For hundreds of years chemistry made little headway. Then, in 1525, a Swiss doctor and scientist spoke up. He had the imposing name of Theophrastus Bombastus Paracelsus von Hohenheim. He challenged his students to tear up their books with the old theories that had been developed through reasoning only and to find out for themselves *through experiments* whether a scientific theory was right or wrong. But only a few people paid attention to him.

More than a hundred years passed before an Englishman, Robert Boyle, in 1661, succeeded in killing off the old idea of the four elements. He did it by establishing that there are many elements — substances that cannot be formed by other substances and cannot be broken into other substances.

Another hundred years went by. Then, at the time of the American Revolution, the day finally dawned for modern chemistry.

A Swede, Karl Scheele, and an Englishman, Joseph Priestley, discovered oxygen, and a Frenchman, Antoine Laurent Lavoisier, explained the true nature of burning and made up the first scientific listing of all known elements — twenty-eight at the time.

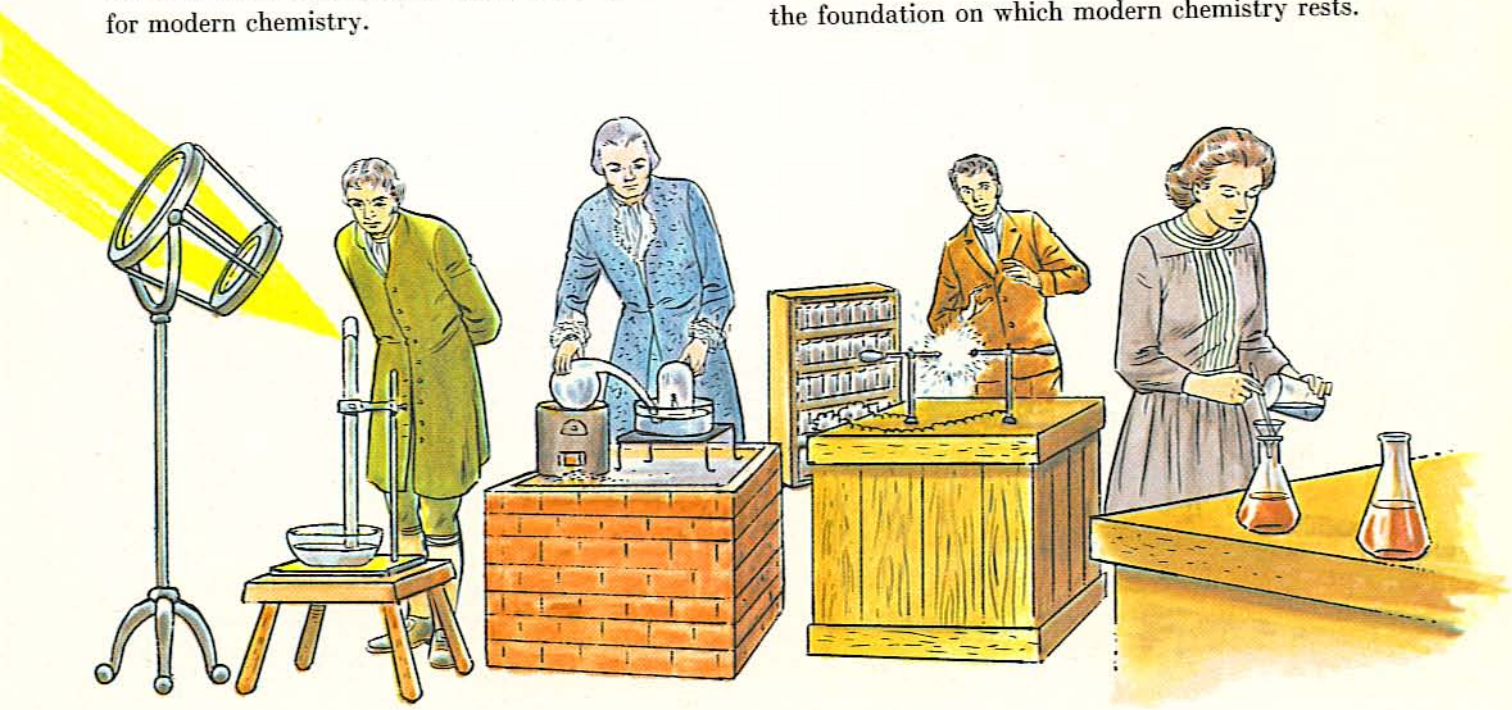
Within a few years, more elements were found. With the help of electricity, an English chemist, Humphry Davy, in a single year brought to light six new metals — among them sodium, potassium, calcium, and magnesium.

Twenty years later, in 1828, another important break-through occurred. A German chemist, Friedrich Wohler, working in his laboratory produced a chemical, urea, that had never before been made outside the body of a living animal.

More and more things were happening. New elements were discovered, new chemicals created. The advances in chemistry greatly influenced industry, agriculture and medicine.

And then, in 1898, the Polish-born Marie Curie and her French husband, Pierre, discovered the "miracle element," radium. This opened up a whole new age in chemistry.

Within the last fifty years, chemistry has moved forward with giant steps. But not a single one of these steps would have been possible without the dedicated work of the chemists of the past who laid the foundation on which modern chemistry rests.



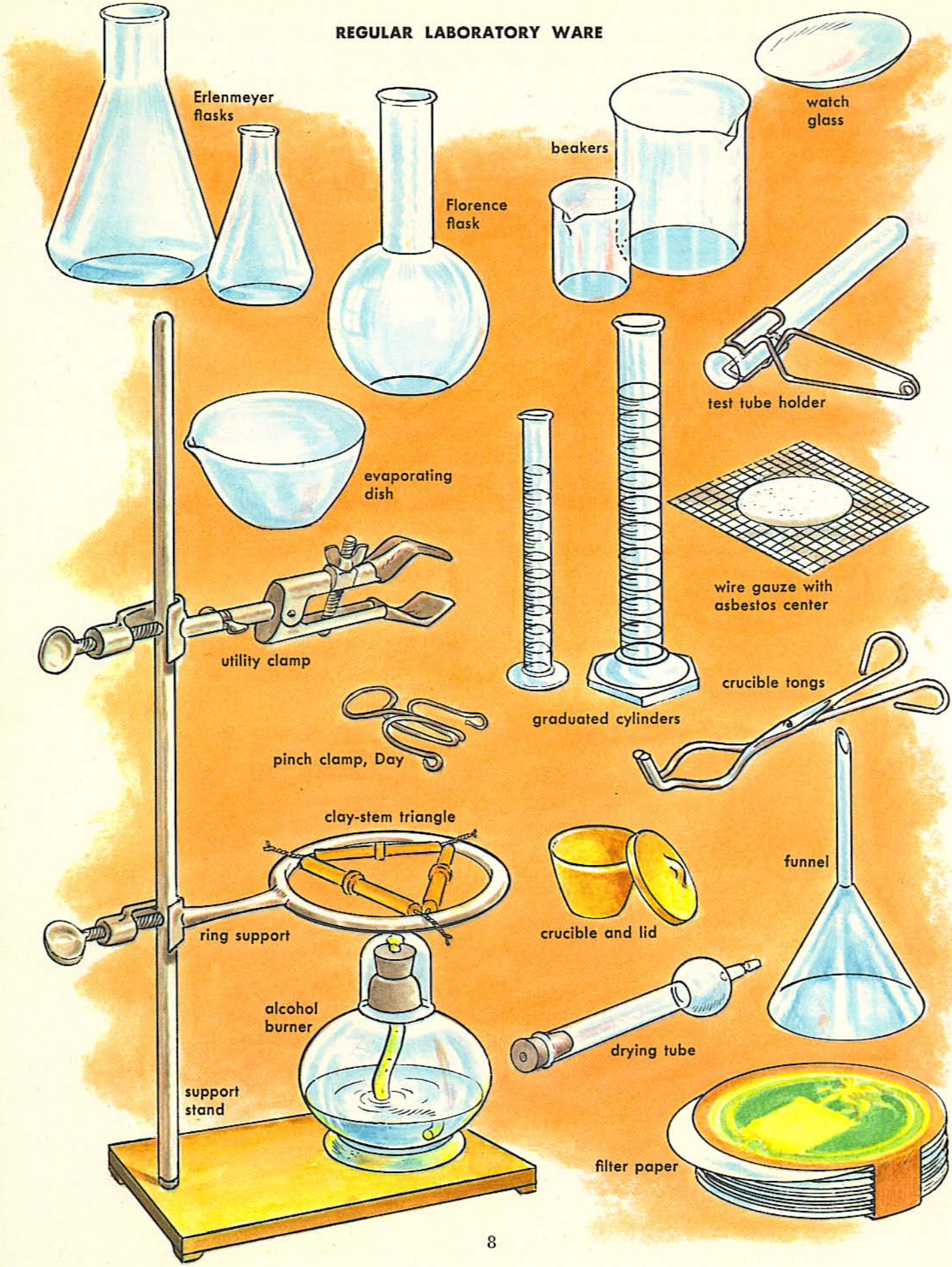
PRIESTLEY USED HEAT OF SUN TO PRODUCE OXYGEN.

LAVOISIER GAVE THE RIGHT EXPLANATION OF BURNING.

DAVY BROUGHT ELECTRICITY INTO CHEMICAL RESEARCH.

MARIE CURIE AND HER HUSBAND DISCOVERED RADIUM.

REGULAR LABORATORY WARE



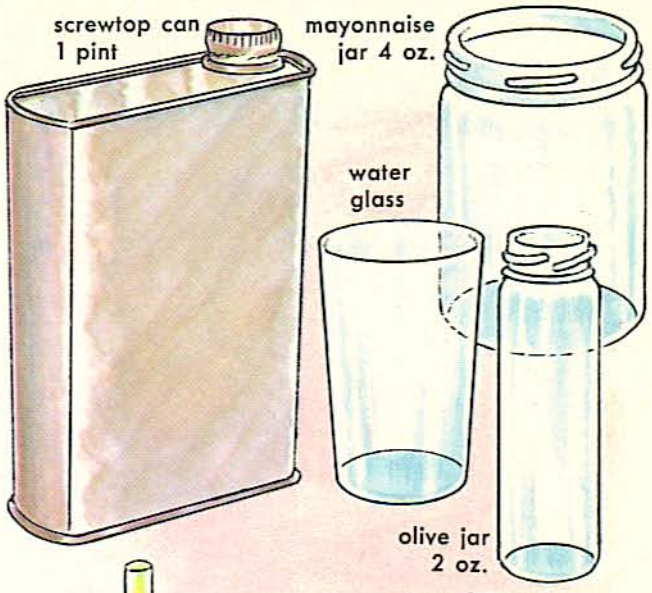
Equipment for Chemistry

Some of the greatest discoveries in chemistry were made by scientists who had no special equipment but simply used whatever was at hand.

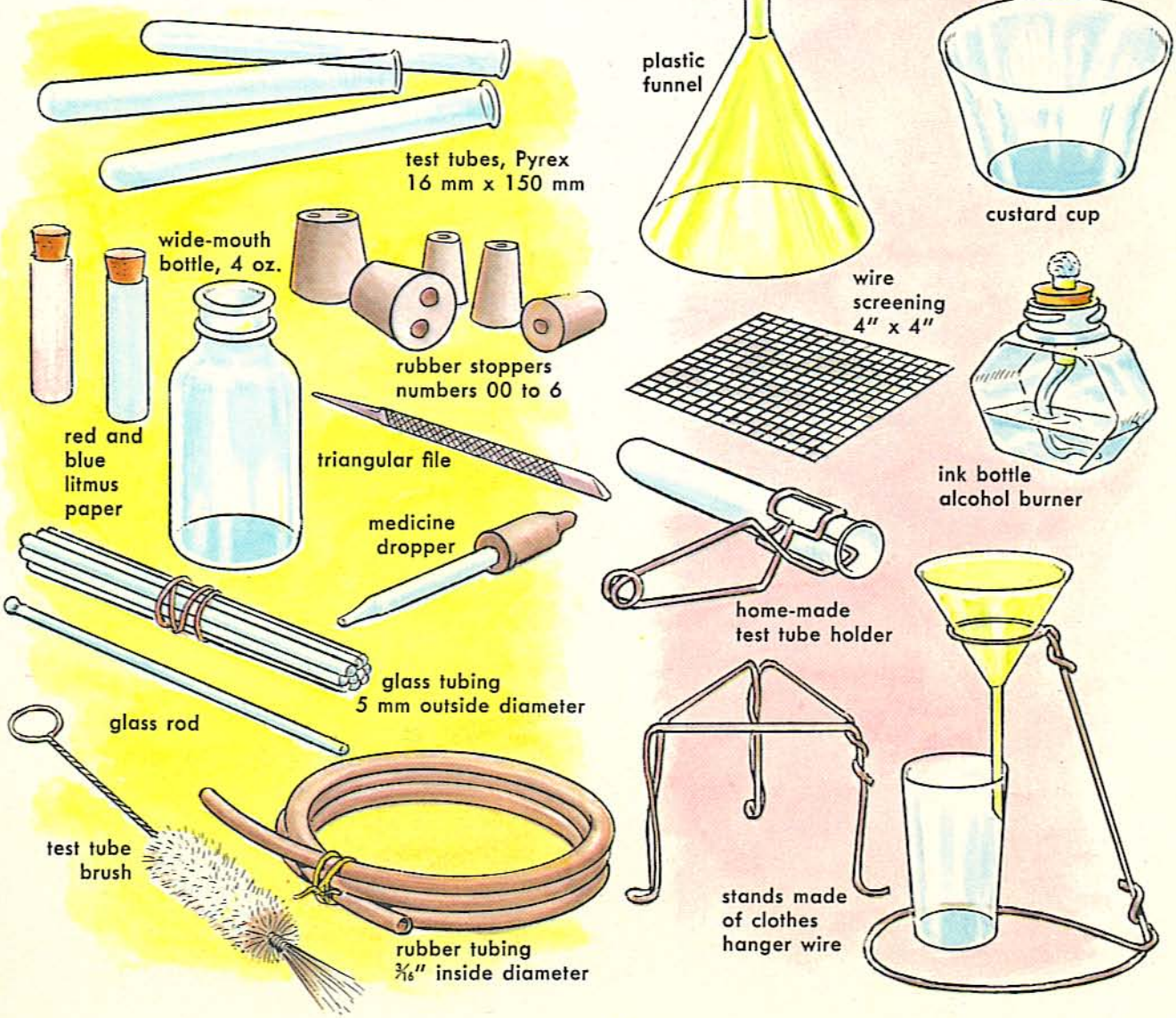
In your home lab experiments it will pay you to follow the example of these early chemists. Put your imagination to work. Use whatever suitable equipment you can find around the house (as suggested in column to the right) and buy only what is absolutely necessary (as shown below). Some items may be purchased in a local drugstore or scientific supply shop. If not, you can buy them from one of the suppliers listed on page 110.

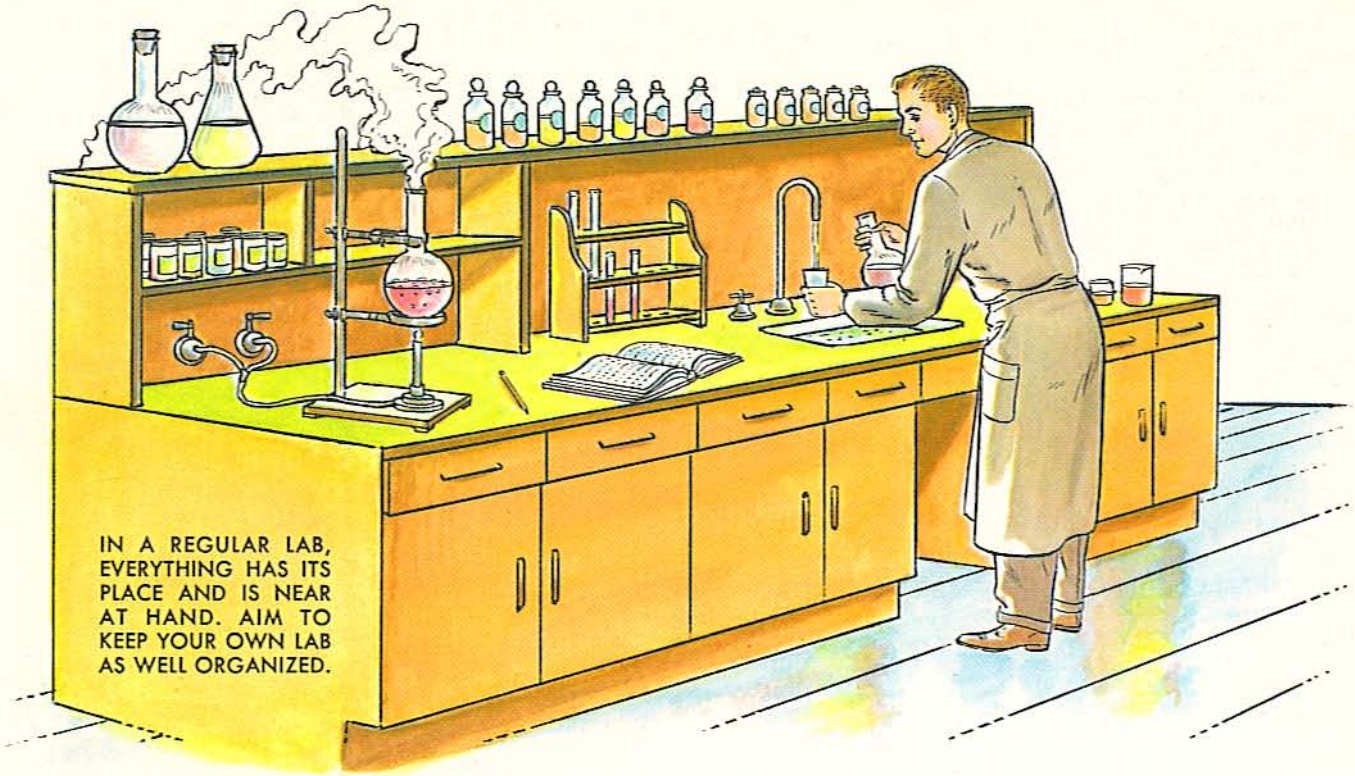
Later on — if you really get excited about chemistry — you may want to use your pocket money for some of the lab equipment shown on page 8.

IMPROVISED EQUIPMENT FOR HOME LAB

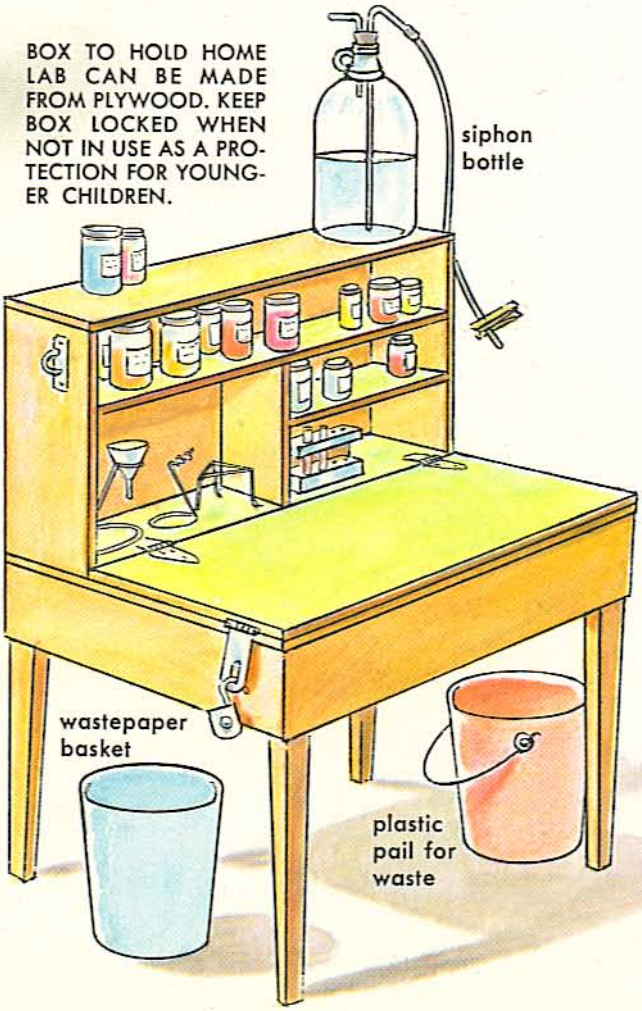


LABORATORY WARE FOR HOME LAB





Setting Up Your Home Laboratory



It is possible that you may be permitted to work at the kitchen table when this is not in use. But it is far better if you have a place where you will not be disturbed and where you can store your equipment — a corner in your room, or in the basement or the garage.

These are the things you'll need in your lab:

Work Table. An old, sturdy table will do. Cover it with a plastic top to protect the wood.

Water Supply. If you have a faucet nearby, fine. Otherwise, make a siphon bottle (page 11).

Waste Disposal. If you can dump your waste directly into the kitchen *drain* (NOT into the sink), you are all right. If not, collect it in a plastic pail to be thrown out when you're finished.

Source of Heat. In the regular laboratory, special gas burners are used. In the home lab, you can use a burner for denatured alcohol. Have a shallow metal pan under the burner for fire safety.

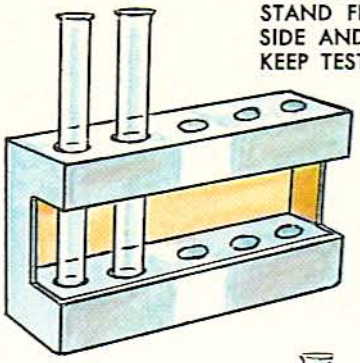
Storage. If there's no one around to disturb your chemicals and equipment, an open shelf is OK. Otherwise, use a box that can be locked up.

Containers. Keep chemicals in glass jars and bottles. LABEL THEM ALL CLEARLY.

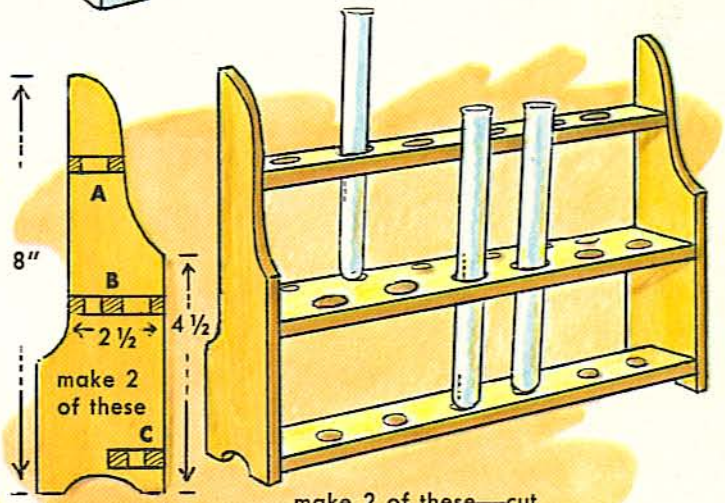
Stands. Make your own test tube stand as well as stands for holding glassware for heating.

IF YOU DO NOT HAVE RUNNING WATER IN YOUR LAB, MAKE SIPHON FROM A 1-GALLON BOTTLE. START SIPHON BY BLOWING THROUGH L-SHAPED GLASS TUBE. SPRING CLOTHES PIN REGULATES WATER FLOW.

YOU CAN MAKE A SIMPLE TEST TUBE STAND FROM A SALT BOX. CUT OUT SIDE AND USE CUT-OUT SECTION TO KEEP TEST TUBES LINED UP STRAIGHT.



REGULAR LAB STAND FOR TEST TUBES CAN BE PRODUCED FROM STRIPS OF ¼-INCH WOOD. FOLLOW PATTERNS BELOW TO MAKE TWO SIDES AND THREE HORIZONTAL PIECES. DRILL HOLES. ASSEMBLE WITH THIN TACKS.

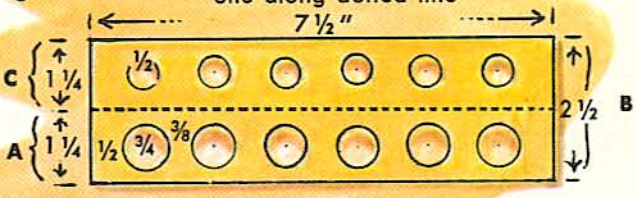


cut lipstick container at dotted line

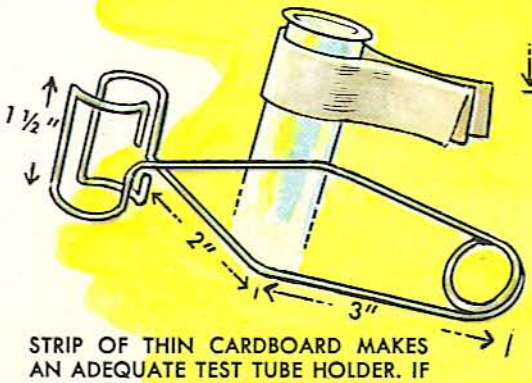
use half of Thermos bottle cork

"screw" lipstick tube through cork

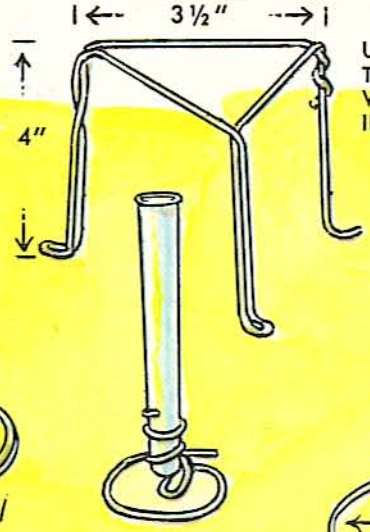
make 2 of these—cut one along dotted line



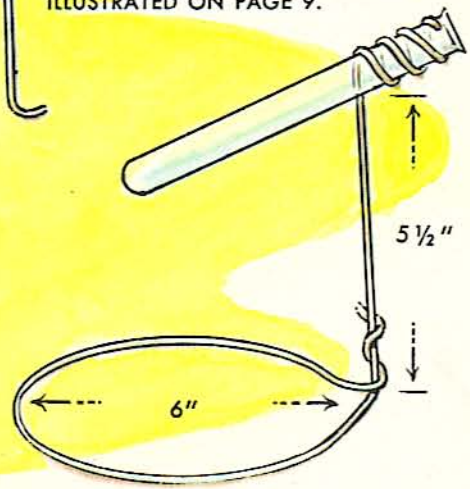
MAKE AN ALCOHOL BURNER FROM INK BOTTLE, THERMOS BOTTLE CORK, AND LIPSTICK CONTAINER. BUY WICK AT A HARDWARE STORE.

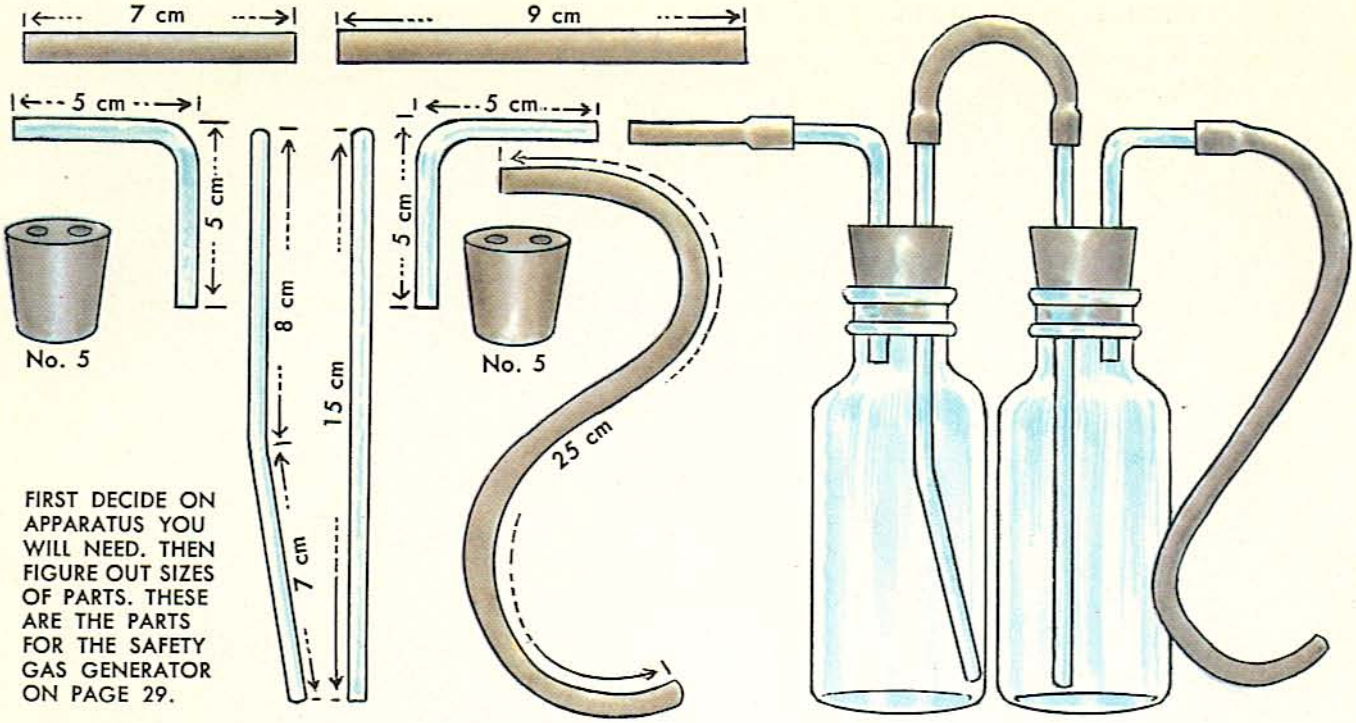


STRIP OF THIN CARDBOARD MAKES AN ADEQUATE TEST TUBE HOLDER. IF YOU LIKE, YOU CAN MAKE A HOLDER FROM CLOTHES HANGER WIRE AS SHOWN HERE.



USE A PAIR OF PLIERS FOR MAKING THESE STANDS FROM CLOTHES HANGER WIRE. ALSO MAKE THE FUNNEL STAND ILLUSTRATED ON PAGE 9.





FIRST DECIDE ON APPARATUS YOU WILL NEED. THEN FIGURE OUT SIZES OF PARTS. THESE ARE THE PARTS FOR THE SAFETY GAS GENERATOR ON PAGE 29.

Making Apparatus for Experiments

MOST of your chemical experiments you will perform in test tubes and jars. But occasionally you will need an apparatus — a device consisting of bottles and stoppers, glass and rubber tubing.

A good chemist takes pride in his apparatus. He makes it with great care — not just for looks but, more important, for safety. An apparatus that leaks flammable gas can be very dangerous.

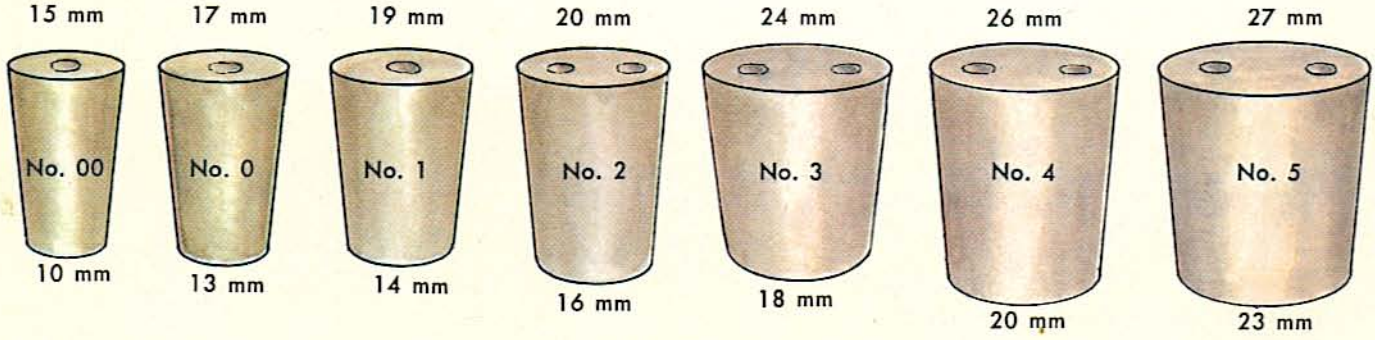
Before you start to put an apparatus together, make a simple drawing of it so that you will know what it will consist of. Then get out the various parts you will need to put it together.

To make an apparatus, you need to know how to cut a glass tube, how to bend it, and how to draw it to a jet point. See page 13.

It is wise to use glass tubes of one diameter only, with rubber tubing to fit. Glass tubes of an *outside* diameter of 6 millimeters fit snugly into the holes in the usual rubber stoppers. Rubber tubing of an *inside* diameter of 3/16" fits over the 6mm glass tubes.

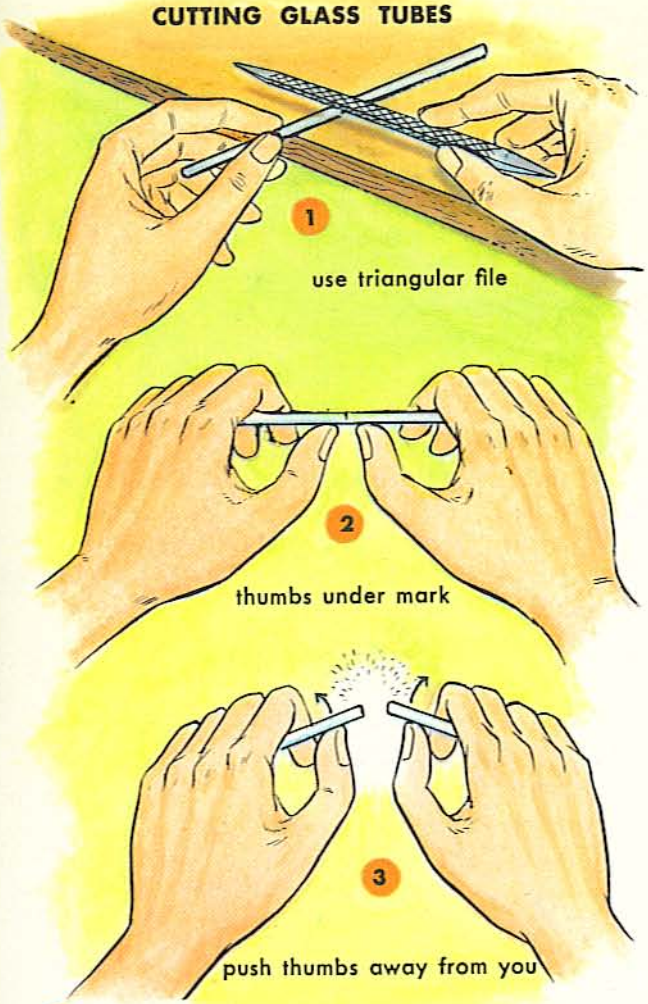
To determine the right size stoppers to use in the bottles of your apparatus, measure the mouths of the bottles against the stoppers shown below in actual size. Order stoppers by number. Keep a selection of different sizes on hand.

Follow the safety precautions on page 16.



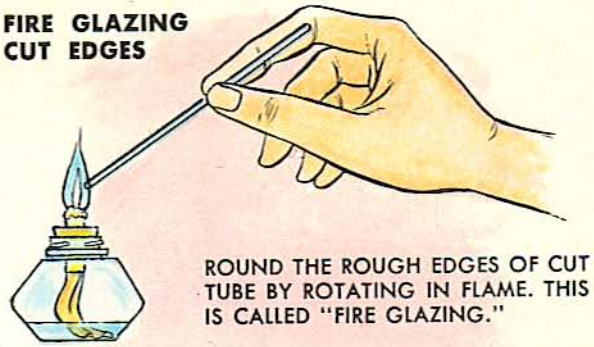
THESE ARE THE ACTUAL SIZES OF RUBBER STOPPERS. BY MEASURING THEM AGAINST YOUR LAB WARE YOU WILL KNOW WHICH TO ORDER. No. 0 FITS THE 16 mm TEST TUBE. No. 5 FITS 4-OZ. WIDE-MOUTH BOTTLE.

CUTTING GLASS TUBES



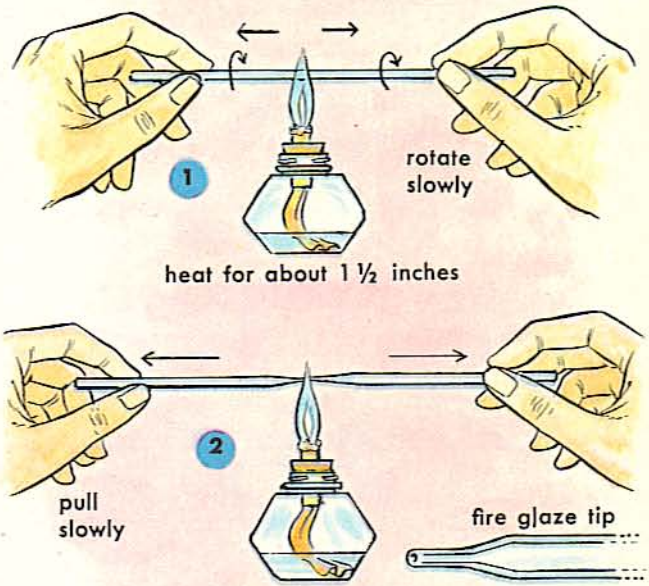
1. MAKE SINGLE SCRATCH WITH FILE—DO NOT "SAW."
2. HOLD TUBE WITH BOTH HANDS, THUMBS BELOW MARK.
3. SNAP TUBE INTO TWO PIECES WITH A QUICK JERK.

FIRE GLAZING CUT EDGES



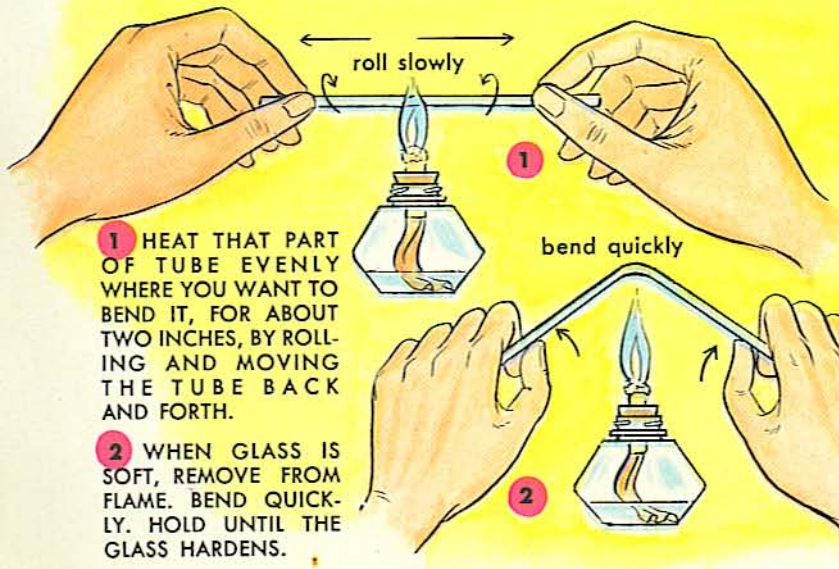
ROUND THE ROUGH EDGES OF CUT TUBE BY ROTATING IN FLAME. THIS IS CALLED "FIRE GLAZING."

MAKING JET POINT

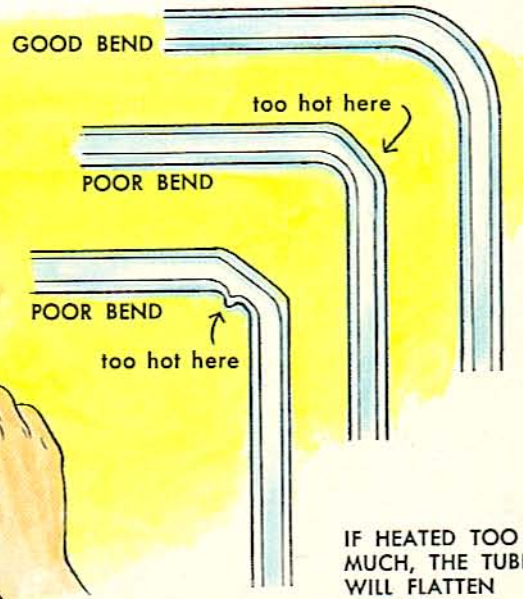


1. HEAT SECTION OF TUBE EVENLY WHERE YOU WANT TO FORM JET POINT, BY ROLLING IT IN THE FLAME.
2. WHEN SOFT, PULL SLOWLY. CUT APART WHEN HARD.

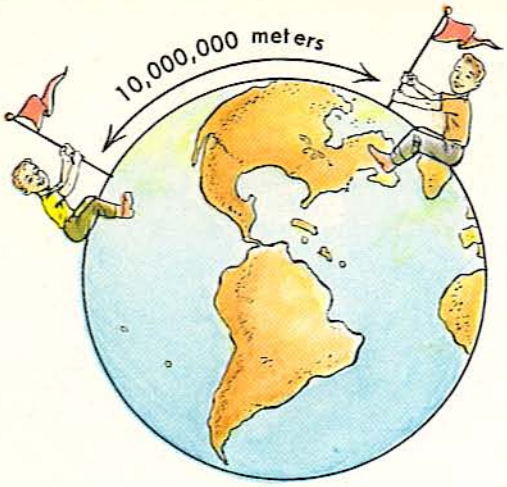
BENDING GLASS TUBES



1. HEAT THAT PART OF TUBE EVENLY WHERE YOU WANT TO BEND IT, FOR ABOUT TWO INCHES, BY ROLLING AND MOVING THE TUBE BACK AND FORTH.
2. WHEN GLASS IS SOFT, REMOVE FROM FLAME. BEND QUICKLY. HOLD UNTIL THE GLASS HARDENS.



IF HEATED TOO MUCH, THE TUBE WILL FLATTEN OR "BUCKLE."



Scientific Measurements

In science, the metric system is preferred over our usual system. It is much easier to work with when once you have learned it — for instead of dividing or multiplying by 12 or 32 or 16 to go from one unit to the next, you simply move the decimal point. Just remember these two things:

1. That the names of the basic units are meter for lengths, liter for volumes, grams for weights — abbreviated to m, l, and g (without a period after them).

2. That 1000 of a kind are called kilo; 100, hekto; 10, deca; 1/10 is called deci; 1/100, centi; 1/1000, milli.

METRIC UNITS OF LENGTH

1000 meters (m) = 1 kilometer (km)
1 meter (m) = 1000 millimeters (mm)
1 meter (m) = 39.37 inches
2.540 centimeters (cm) = 1 inch

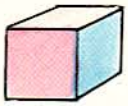
METRIC UNITS OF VOLUME

1 liter (l) = 1000 cubic centimeters (cm^3 or cc)
1 liter (l) = 1000 milliliters (ml)
1 liter (l) = 1.06 quarts (liquid)
0.946 liter (l) = 1 quart (liquid)

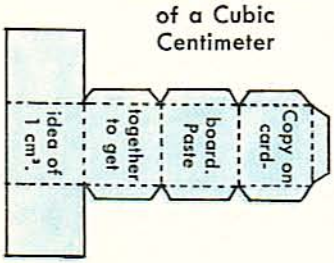
METRIC UNITS OF WEIGHT

1000 grams (g) = 1 kilogram (kg)
1 gram (g) = 1000 milligrams (mg)
1 gram (g) = 0.035 ounces avoirdupois
28.350 grams (g) = 1 ounce avoirdupois

1 cubic centimeter =
1 cm^3 = 1 cc =
1 milliliter =
1 ml



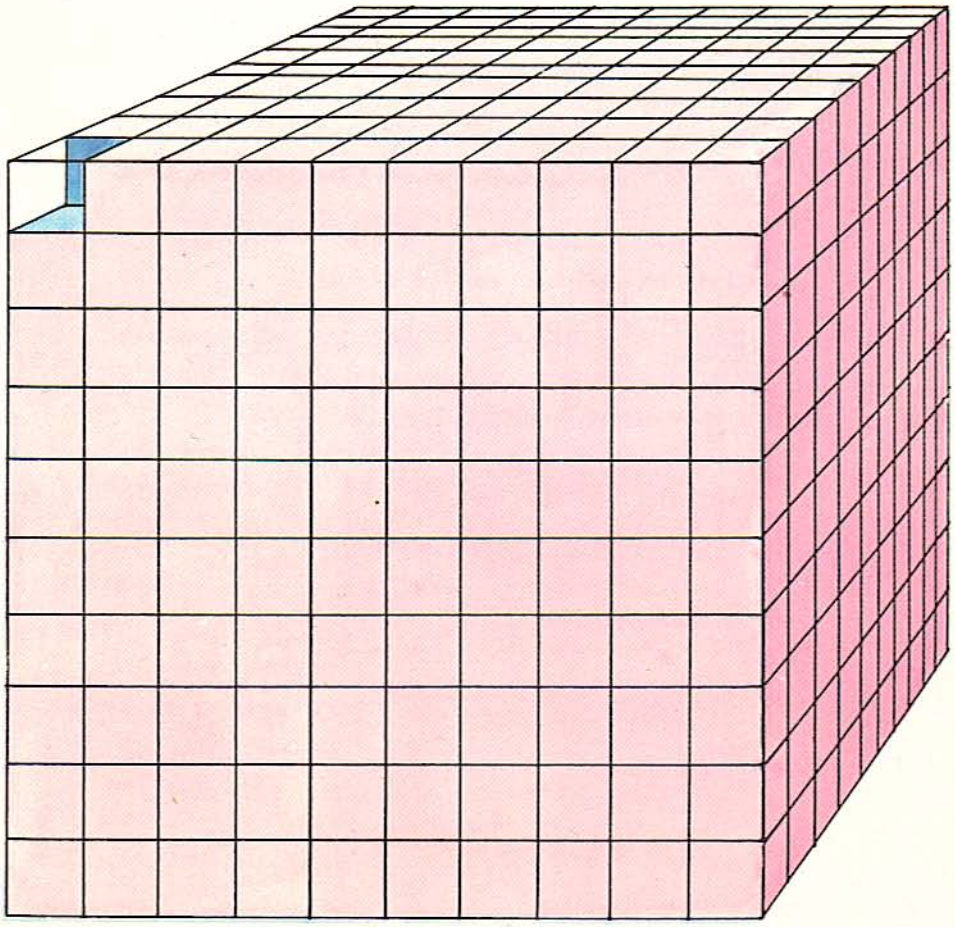
Making a Model of a Cubic Centimeter



BOTTOM LINE OF FIGURE TO THE RIGHT IS 1 DECIMETER (1 dm) OR 10 CENTIMETERS (10 cm) OR 100 MILLIMETERS (100 mm).

FRONT SURFACE OF FIGURE IS 1 SQUARE DECIMETER (1 dm^2) OR 100 SQUARE CENTIMETERS (100 cm^2).

VOLUME OF WHOLE FIGURE IS 1 CUBIC DECIMETER (1 dm^3) OR 1000 CUBIC CENTIMETERS (1000 cm^3 OR 1000 cc) OR 1000 MILLILITERS (1000 ml) OR 1 LITER (1 l). 1 l WATER WEIGHS 1 KILOGRAM (1 kg).



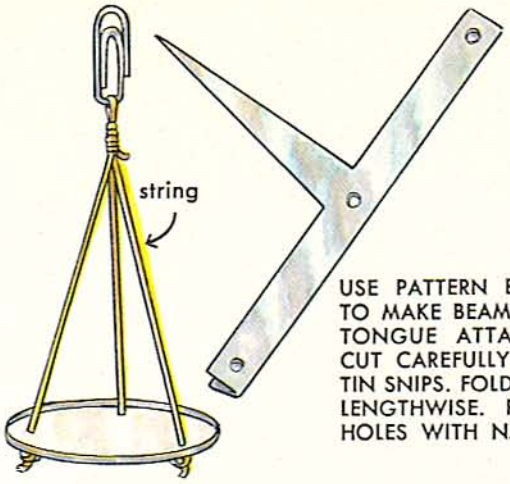
1 cubic decimeter = 1 liter = 1000 cubic centimeters = 1000 milliliters

centimeter ruler

inch ruler

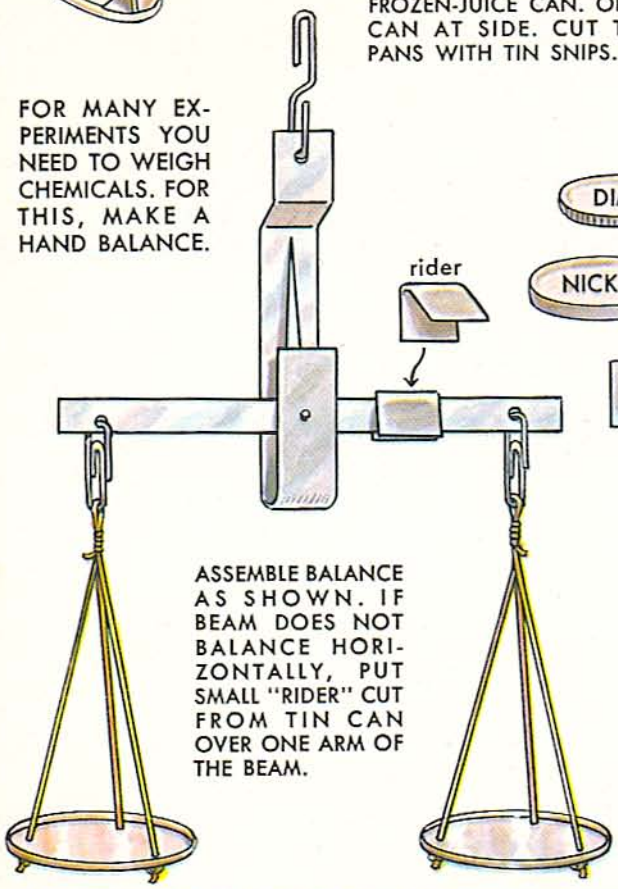
MAKING A HAND BALANCE

MAKE THE TWO PANS FOR THE HAND BALANCE FROM TOP AND BOTTOM OF A FROZEN-JUICE CAN. OPEN CAN AT SIDE. CUT THE PANS WITH TIN SNIPS.

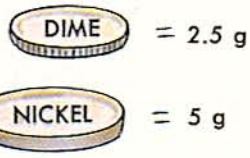


USE PATTERN BELOW TO MAKE BEAM WITH TONGUE ATTACHED. CUT CAREFULLY WITH TIN SNIPS. FOLD BEAM LENGTHWISE. PUNCH HOLES WITH NAIL.

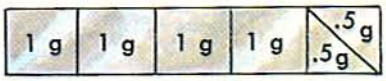
FOR MANY EXPERIMENTS YOU NEED TO WEIGH CHEMICALS. FOR THIS, MAKE A HAND BALANCE.



ASSEMBLE BALANCE AS SHOWN. IF BEAM DOES NOT BALANCE HORIZONTALLY, PUT SMALL "RIDER" CUT FROM TIN CAN OVER ONE ARM OF THE BEAM.



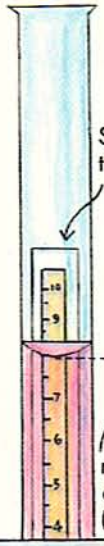
YOU CAN USE COINS TO WEIGH 2.5 g, 5 g, 7.5 g, 10 g, ETC. FOR SMALLER WEIGHTS, CUT A TIN STRIP TO WEIGH THE SAME AS A NICKEL. THEN CUT IT IN FOUR 1 g AND TWO .5 g WEIGHTS.

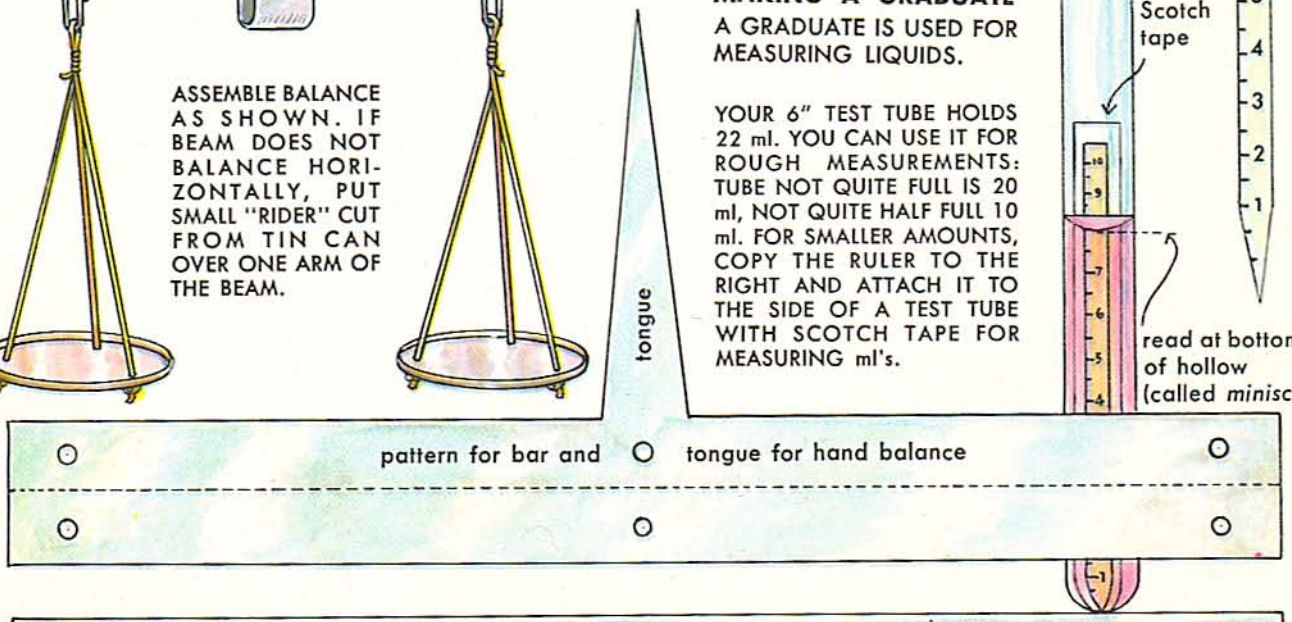


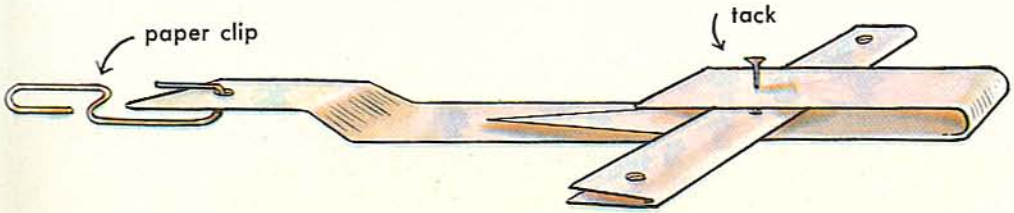
MAKING A GRADUATE

A GRADUATE IS USED FOR MEASURING LIQUIDS.

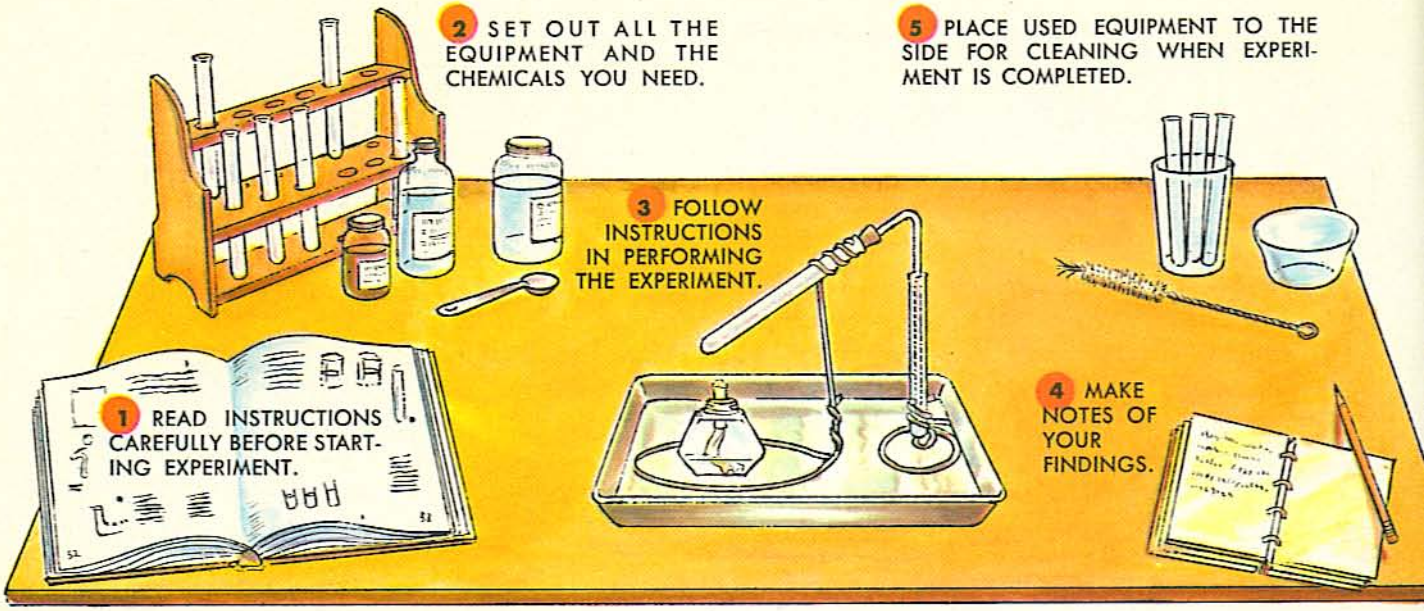
YOUR 6" TEST TUBE HOLDS 22 ml. YOU CAN USE IT FOR ROUGH MEASUREMENTS: TUBE NOT QUITE FULL IS 20 ml, NOT QUITE HALF FULL 10 ml. FOR SMALLER AMOUNTS, COPY THE RULER TO THE RIGHT AND ATTACH IT TO THE SIDE OF A TEST TUBE WITH SCOTCH TAPE FOR MEASURING ml's.







COPY THE PATTERNS ABOVE ONTO A PIECE OF PAPER. TRANSFER DESIGNS TO PIECE OF TIN CAN. CUT OUT AND BEND AS SHOWN TO THE LEFT.



Correct Laboratory Techniques

In your home laboratory, three considerations are of the greatest importance: SAFETY, NEATNESS, and EXACTNESS.

SAFETY — All the experiments in this book are safe when done in the correct laboratory way as shown on these pages.

Treat chemicals with respect. Never taste anything unless specifically told to do so. If there are younger children in the family, lock up your chemicals when you are not working with them.

Protect your clothes with a plastic apron.

Be careful with fire. When you use your alcohol burner, have a metal pan under it for safety.

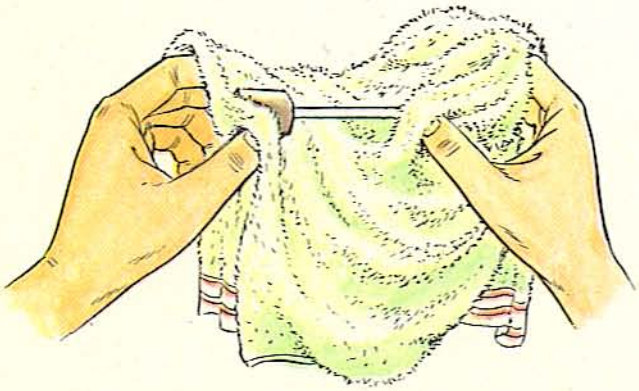
NEATNESS — Get the habit of lining up equipment and chemicals you need on one side and placing used items on the opposite side — keeping the space between them clear for your experiments.

Put chemicals away and clean glassware as soon as you have finished an experiment.

EXACTNESS — Label all bottles and jars containing chemicals clearly and correctly.

Where amounts of chemicals are not given, use the smallest amount that will tell you what you want to know.

Observe the chemical reactions carefully and make complete notes of them as you go along.



PLAY SAFE WHEN YOU PUT A GLASS TUBE IN A STOPPER. PROTECT YOUR HANDS BY WRAPPING TOWEL AROUND THEM. MOISTEN GLASS TUBE AND STOPPER WITH WATER, THEN PUSH THE TUBE INTO THE STOPPER WITH A SCREW-DRIVER MOTION.

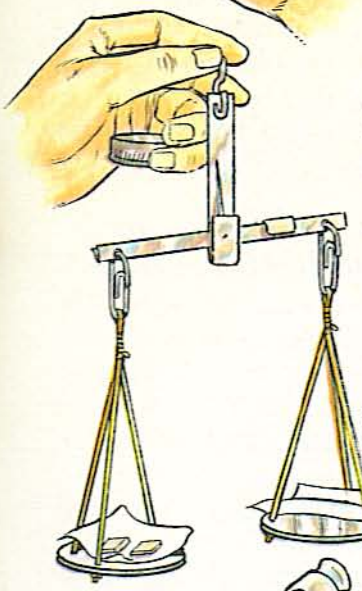
WHEN YOU MAKE AN APPARATUS FOR A CHEMICAL EXPERIMENT, MAKE SURE THAT ALL CONNECTIONS ARE AIR-TIGHT. USE THE RIGHT SIZE STOPPER FOR MOUTH OF CONTAINER, GLASS TUBES THAT FIT SNUGLY INTO THE STOPPER HOLES, AND TIGHT-FITTING RUBBER TUBING.

MAKE IT A HABIT TO READ A LABEL TWICE TO BE SURE YOU HAVE THE RIGHT CHEMICAL.

TRAINED CHEMISTS NEVER PUT STOPPER OF BOTTLE ON DESK—THEY KEEP IT IN ONE HAND.

1 TAKE STOPPER OFF BOTTLE WITH YOUR LEFT HAND.

2 KEEP STOPPER IN THE LEFT HAND WHILE YOU POUR FROM THE BOTTLE IN YOUR RIGHT, WITH THE LABEL FACING UP. THEN REPLACE THE STOPPER.

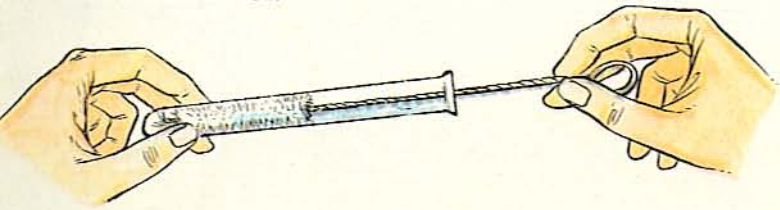


IN WEIGHING OUT A DRY CHEMICAL, PLACE EQUAL-SIZED TISSUE PAPERS ON EACH PAN OF THE SCALE. ROLL JAR BETWEEN FINGERS OR TAP IT GENTLY WITH YOUR INDEX FINGER.

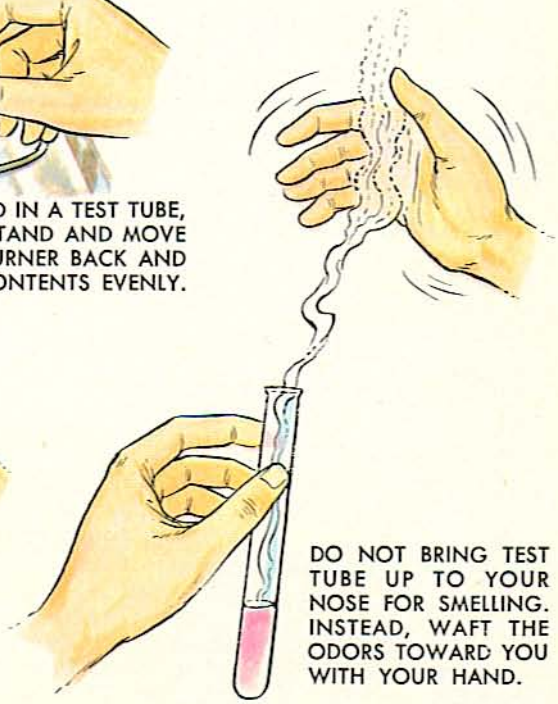
USE A GLASS ROD TO DIRECT THE STREAM WHEN POURING A LIQUID FROM ONE CONTAINER TO ANOTHER.

WHEN HEATING A LIQUID IN A TEST TUBE, HOLD TUBE WITH A HOLDER. KEEP THE TUBE MOVING. DO NOT HEAT BOTTOM OF TUBE—CONTENTS MAY "BUMP" AND SQUIRT OUT. NEVER POINT MOUTH OF TEST TUBE TOWARD YOURSELF OR TOWARD ANYONE ELSE.

WHEN HEATING A SOLID IN A TEST TUBE, PLACE THE TUBE IN A STAND AND MOVE THE FLAME OF YOUR BURNER BACK AND FORTH TO HEAT THE CONTENTS EVENLY.



USE A TEST TUBE BRUSH FOR CLEANING TEST TUBES. RINSE IN COLD WATER.



DO NOT BRING TEST TUBE UP TO YOUR NOSE FOR SMELLING. INSTEAD, WAFT THE ODORS TOWARD YOU WITH YOUR HAND.

Mr. Faraday's Candle

In the winter of 1859, Michael Faraday, a great British scientist, gave a number of lectures for young people. The talks dealt with one subject only: the features or "phenomena" of — a candle!

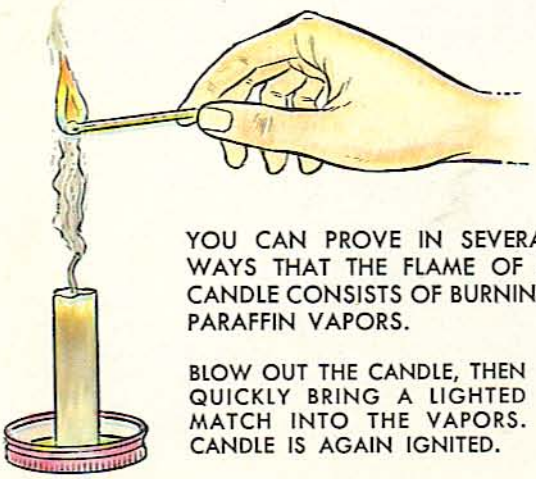
"There is not a law," Faraday told his listeners, "under which any part of this universe is governed which does not come into play and is touched upon in these phenomena. There is no better, there is no more open door by which you can enter into the study of natural philosophy than by considering the phenomena of a candle." He then set out to prove his point by lighting a candle and demonstrating all the processes involved.

In burning a candle you start with a SOLID substance that turns, first, into a LIQUID, then into a GAS (or, more correctly, into a gas-like vapor). The melted candle grease is held in a level position by GRAVITY yet seems to defy gravity by rising in the wick by a force called CAPILLARY ACTION. In burning, the candle produces ENERGY in the form of LIGHT and HEAT. At the same time, it goes into CHEMICAL REACTIONS that reveal what it is made of.

As you enter the study of chemistry, you can do no better than to repeat for yourself some of the experiments that Mr. Faraday demonstrated to his young audience.

PLACE A BURNING CANDLE IN THE SUN AND CATCH THE SHADOW ON A PIECE OF WHITE PAPER. YOU WILL DISCOVER THAT IT IS THE BRIGHTEST PART OF THE FLAME THAT CASTS THE DARKEST SHADOW.

CANDLE FLAME IS BURNING VAPOR

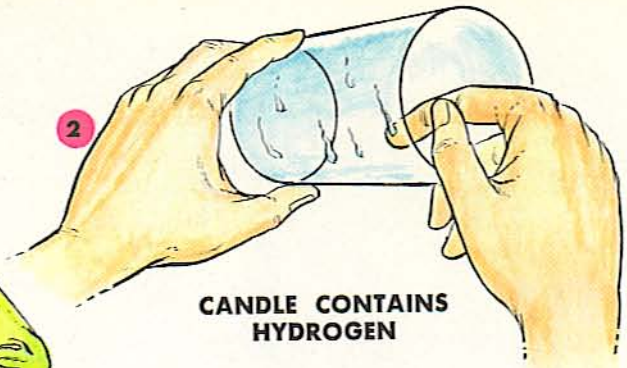


YOU CAN PROVE IN SEVERAL WAYS THAT THE FLAME OF A CANDLE CONSISTS OF BURNING PARAFFIN VAPORS.

BLOW OUT THE CANDLE, THEN QUICKLY BRING A LIGHTED MATCH INTO THE VAPORS. CANDLE IS AGAIN IGNITED.

MAKE A GAS WORKS FROM A CANDLE BY HOLDING A GLASS TUBE, DRAWN TO A POINT, IN THE FLAME AND LIGHTING THE VAPORS AT THE END OF TUBE.

YOU CAN ACTUALLY LEAD THE VAPORS FROM A BURNING CANDLE THROUGH A BENT GLASS TUBE INTO A WATER GLASS. IF LEFT ALONE, WHITISH VAPORS CONDENSE INTO A SOLID.

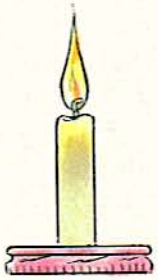


CANDLE CONTAINS HYDROGEN

A CANDLE FORMS WATER WHEN IT BURNS. ONLY HYDROGEN BURNING IN OXYGEN FORMS WATER.

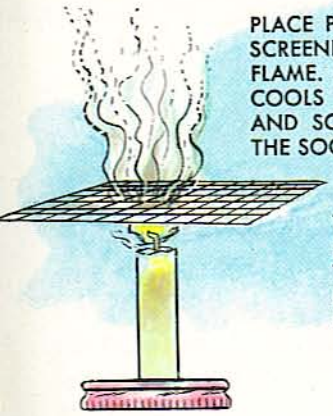
1 HOLD A COLD GLASS FOR A FEW MOMENTS OVER A BURNING CANDLE. DEW FORMS ON THE INSIDE.

2 BY WIPING FINGER INSIDE THE GLASS YOU CAN MAKE DEW FLOW TOGETHER INTO WATER DROPS.

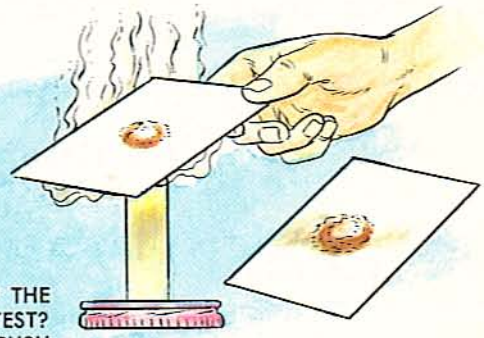


CRUSH A FEW ICE CUBES AND SPRINKLE THEM WITH SALT. WRAP IN ALUMINUM FOIL. HOLD OVER BURNING CANDLE. WATER DROPS WILL FORM ON THE OUTSIDE OF THE FOIL.

CANDLE CONTAINS CARBON

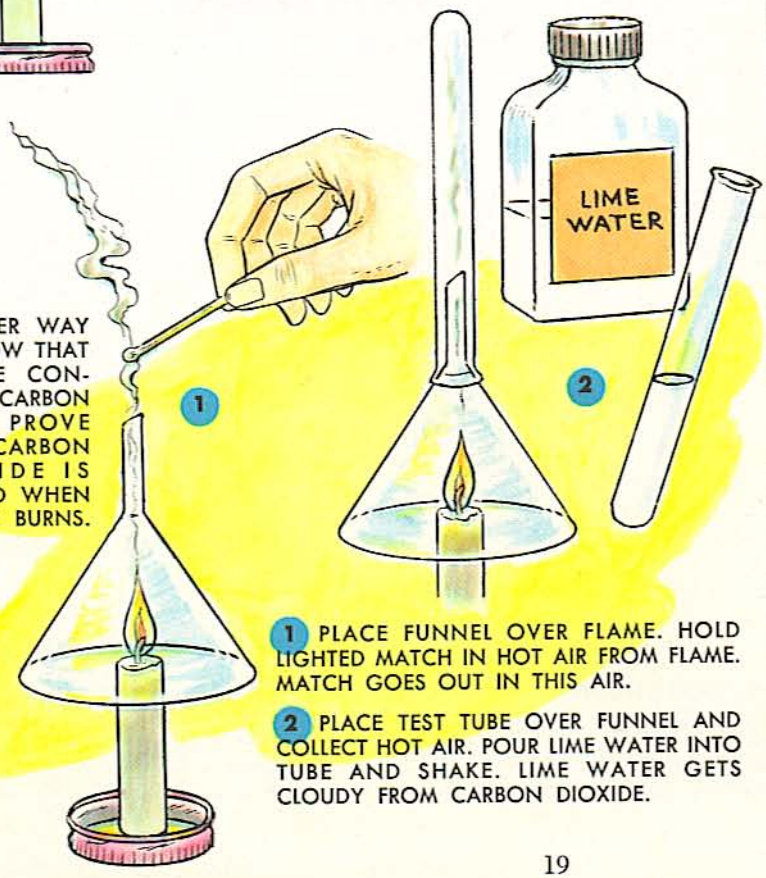


PLACE PIECE OF WIRE SCREENING OVER FLAME. SCREENING COOLS THE FLAME AND SOOT FORMS. THE SOOT IS CARBON.



WHAT PART OF THE FLAME IS HOTTEST? TO FIND OUT, PUSH PIECE OF CARDBOARD SIDEWAYS INTO FLAME. OUTSIDE OF FLAME SCORCHES A SOOTY RING.

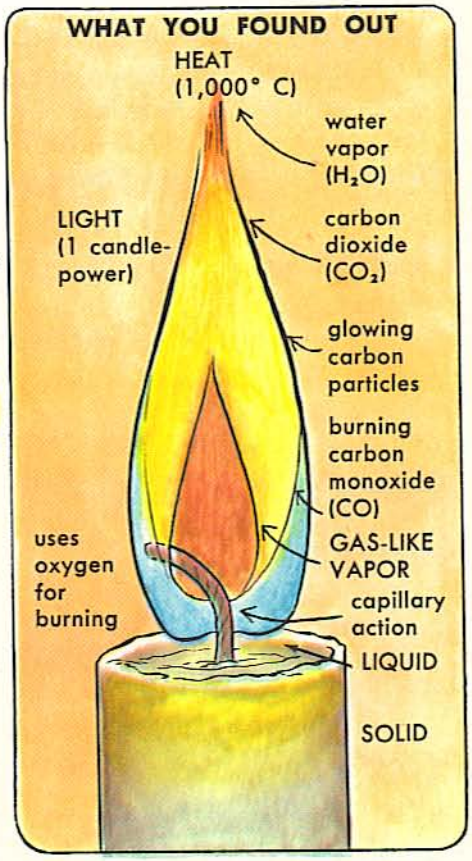
ANOTHER WAY TO SHOW THAT CANDLE CONTAINS CARBON IS TO PROVE THAT CARBON DIOXIDE IS FORMED WHEN CANDLE BURNS.

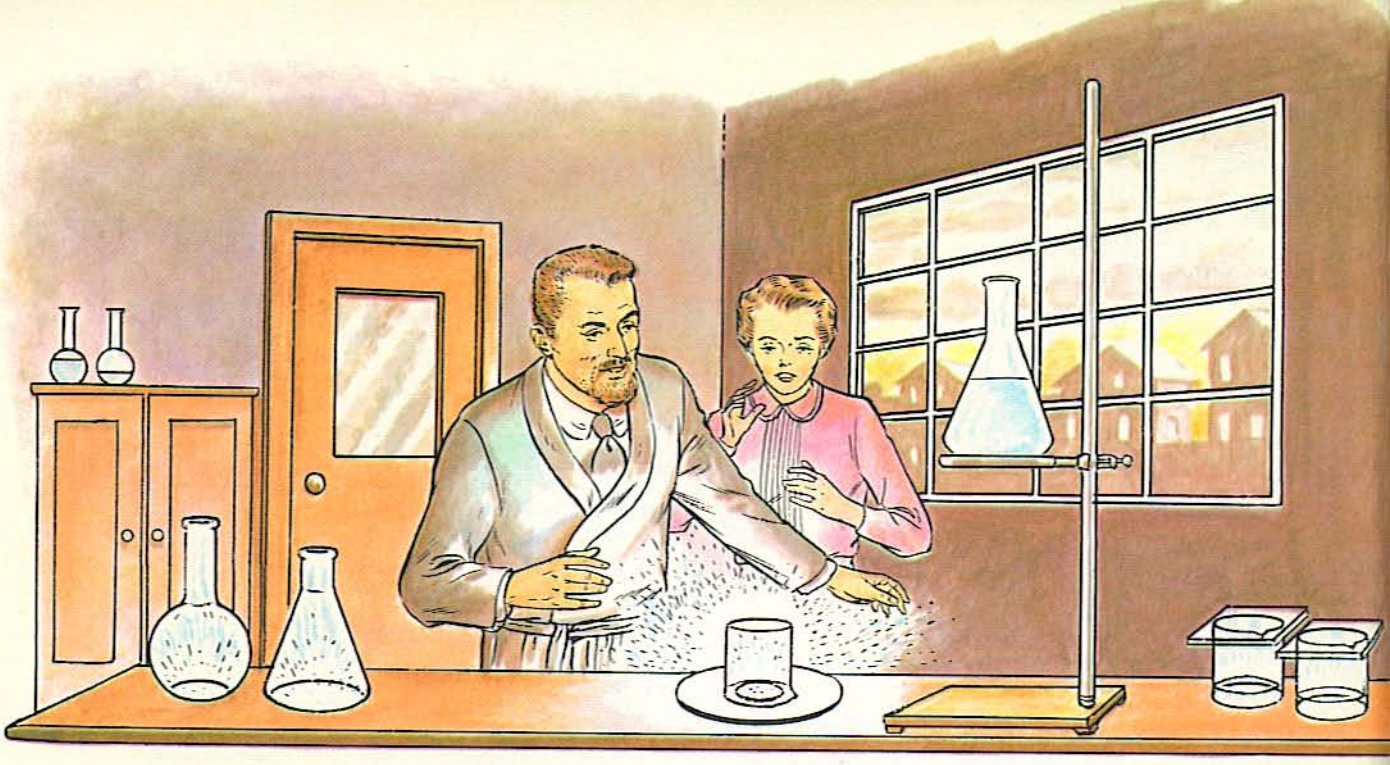


1 PLACE FUNNEL OVER FLAME. HOLD LIGHTED MATCH IN HOT AIR FROM FLAME. MATCH GOES OUT IN THIS AIR.

2 PLACE TEST TUBE OVER FUNNEL AND COLLECT HOT AIR. POUR LIME WATER INTO TUBE AND SHAKE. LIME WATER GETS CLOUDY FROM CARBON DIOXIDE.

WHAT YOU FOUND OUT



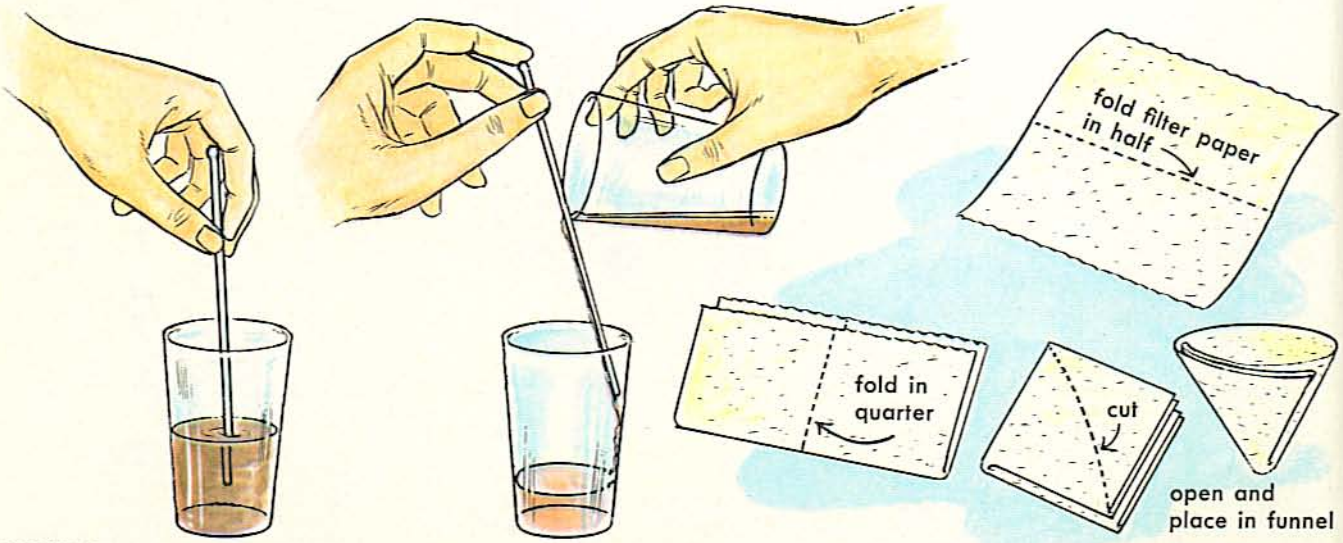


You—Scientist!

In 1896, a young Polish chemist, Marie Curie, and her French husband, Pierre, decided to find out why a certain uranium ore called pitchblende gave off rays that were much stronger than the uranium content of the ore could explain.

They secured a whole ton of powdered ore from a mine in northern Bohemia and set to work. First the powder had to be boiled with strong acids to extract the mysterious substance hidden in it. Then the solution had to be filtered and boiled down. What remained had to be purified by various processes which the Curies had to invent themselves.

After two years of back-breaking work they reached their goal. One night they went to the shed in which they had been working. They opened the door and stepped in without putting on the lights. All around them, the containers that held the solutions of the new substance glowed in the dark! They had discovered a new element — radium — a million times more active than uranium.



SOLUTION—STIR WATER INTO THE MIXTURE OF SALT AND DIRT. THE WATER WILL DISSOLVE THE SALT BUT NOT THE DIRT. YOU NOW HAVE THE SALT IN "WATERY SOLUTION."

DECANTATION—LET DIRT-MIXED SALT SOLUTION STAND UNTIL MOST OF THE DIRT HAS SETTLED. THEN POUR OFF THE LIQUID. THIS PROCESS IS CALLED "DECANTATION."

FILTRATION 1—THE LIQUID IS PROBABLY STILL MURKY. TO CLEAR IT, IT NEEDS TO BE FILTERED BY LETTING IT RUN THROUGH FILTER PAPER (PAPER TOWELING WILL DO).

Why tell again the story of the discovery of radium? Because it contains all the features that show the nature of the true scientist.

Curiosity first. The Curies were curious about the mystery that lay in that greyish-black powder. They became obsessed with a desire to find out — not in the hope of gaining money or fame but to establish a scientific truth.

Before starting their work, the Curies gathered all the known facts about the material with which they were to work. To this knowledge they added their own imagination, figuring out the method they had to use to arrive at the result they were seeking.

For the next two years they literally slaved in the drafty shed that was their laboratory.

After they had made their discovery, the Curies made their method of extracting radium known to the world so that other scientists could check and test what they had done.

As an example of the scientific method the Curies used, let us follow in their footsteps — but with a much simpler problem:

1 MIX THOROUGHLY ONE TABLESPOON OF DIRT AND ONE TEASPOON OF ORDINARY TABLE SALT. NOW DECIDE THAT YOU WANT TO EXTRACT THE SALT FROM THIS MIXTURE AS EARNESTLY AS THE CURIES DECIDED TO EXTRACT THE MYSTERIOUS SUBSTANCE FROM PITCHBLENDE —WITH THE EXCEPTION THAT YOU KNOW WHAT YOU ARE AFTER.

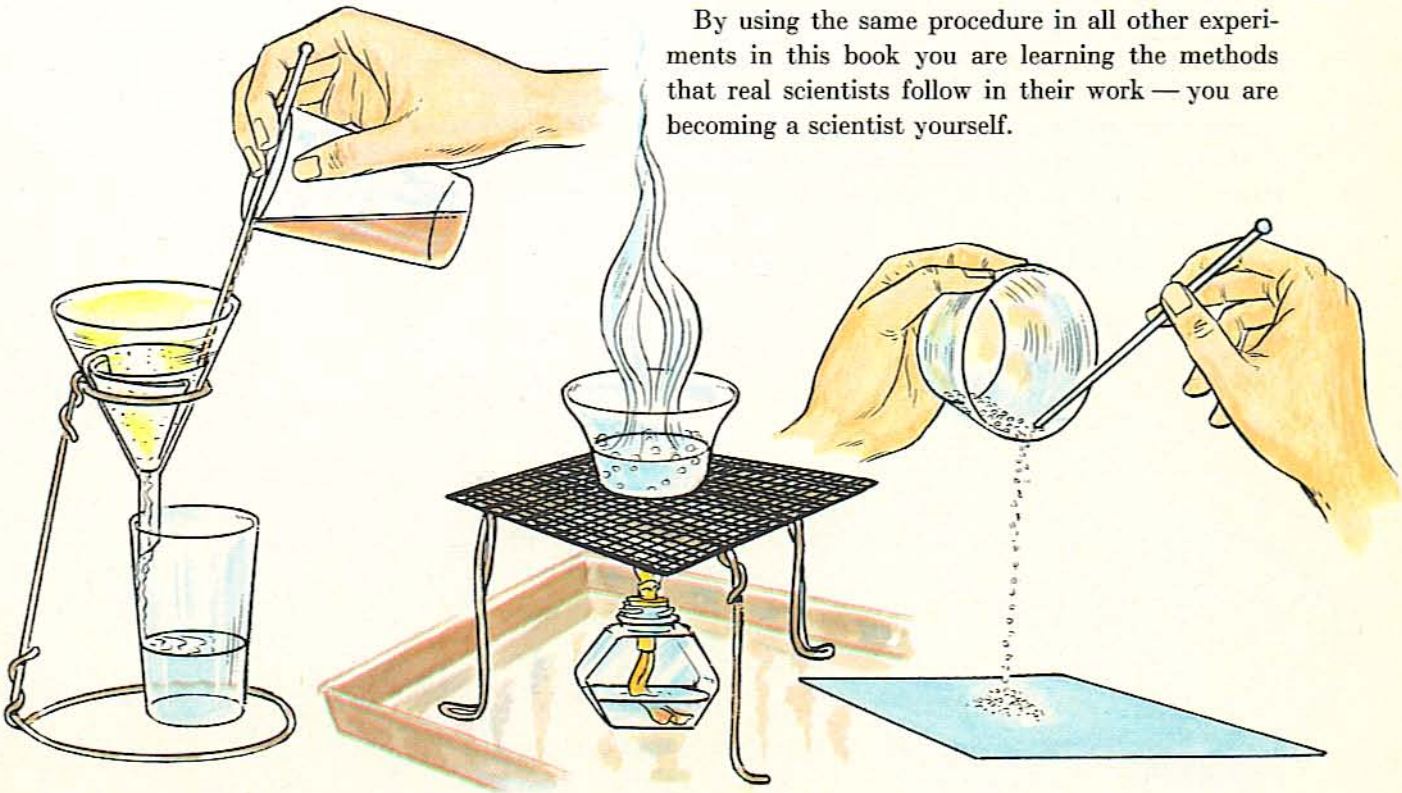
2 GET THE FACTS TOGETHER. DIRT IS "DIRTY," SALT IS WHITE. DIRT PARTICLES ARE OF MANY DIFFERENT SHAPES, SALT CONSISTS OF TINY CUBES. DIRT DOES NOT DISSOLVE IN WATER, SALT DOES.

3 NEXT FIGURE OUT A SUITABLE WAY OF SEPARATING THE TWO SUBSTANCES. ON THE BASIS OF WHAT YOU KNOW YOU SHOULD BE ABLE TO SEPARATE THEM WITH A PAIR OF TINY TWEEZERS—BUT IT WOULD PROBABLY TAKE YOU A YEAR TO DO IT. OR YOU COULD DISSOLVE THE SALT IN WATER AND SEPARATE THE SOLUTION FROM THE INSOLUBLE DIRT.

4 YOU DECIDE ON THE SECOND WAY, USING THE STEPS SHOWN ON THE BOTTOM OF THESE PAGES. IN DOING THIS, YOU DO WHAT THE CURIES DID IN EXTRACTING RADIUM AND LEARN, IN THE PROCESS, THE IMPORTANT LABORATORY TECHNIQUES OF SOLUTION, DECANTATION, FILTRATION, EVAPORATION, AND CRYSTALLIZATION.

5 FINALLY, YOU CHECK THE RESULT. THE WHITE SUBSTANCE LEFT AFTER EVAPORATION SHOULD BE SALT— BUT IS IT? IT LOOKS LIKE SALT. IT TASTES LIKE SALT. BY CHEMICAL TESTS YOU CAN PROVE THAT IT *IS* SALT.

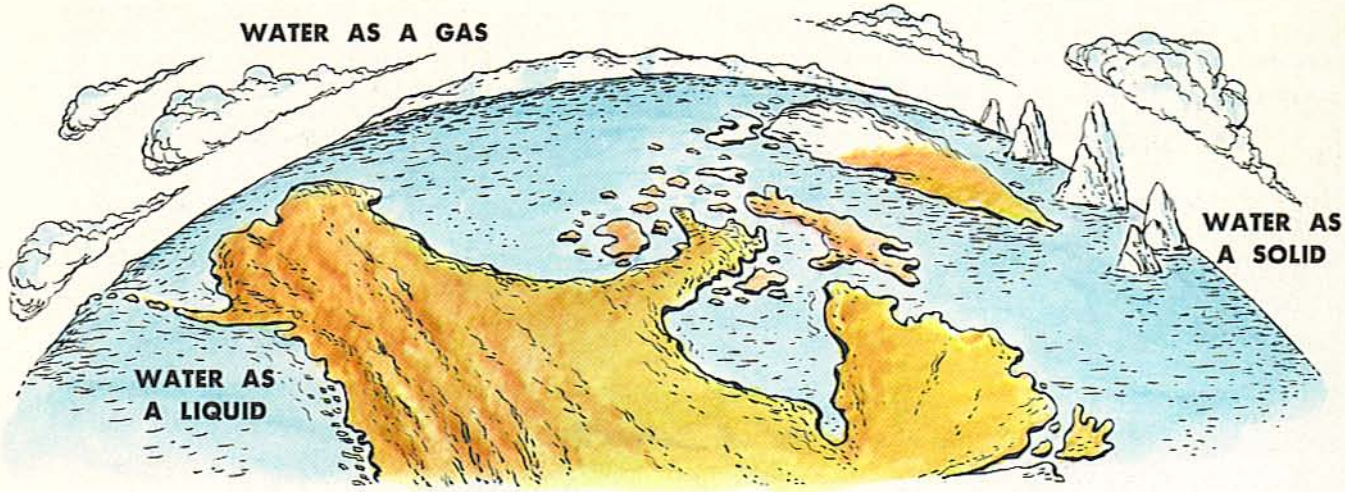
By using the same procedure in all other experiments in this book you are learning the methods that real scientists follow in their work — you are becoming a scientist yourself.



FILTRATION 2—FOLD FILTER PAPER AS SHOWN ON OPPOSITE PAGE AND FIT IT IN FUNNEL. POUR LIQUID ONTO FILTER PAPER. CLEARED LIQUID IS CALLED "FILTRATE."

EVAPORATION—THE FILTRATE CONTAINS THE SALT. THE SALT CAN NOW BE FREED BY REMOVING THE WATER BY BOILING IT AWAY. THIS IS KNOWN AS "EVAPORATION."

CRYSTALLIZATION—AS WATER IS REMOVED, THERE IS TOO LITTLE OF IT LEFT FOR THE SALT TO STAY IN SOLUTION. THE SALT MAKES ITS APPEARANCE AS TINY CRYSTALS.



Elements, Compounds, and Mixtures

IN ALL your experiments in chemistry, you will be dealing with "matter."

Matter is anything that takes up room and has weight (or "mass"). An iron bar is matter — it takes up room and is heavy, as you very well know. Water is matter — it takes up room when you fill a pail with it, and a full pail weighs plenty. The air around you is matter — it takes up lots of room; it may not seem very heavy, yet the earth's atmosphere presses down on every square inch of your body with a weight of almost fifteen pounds.

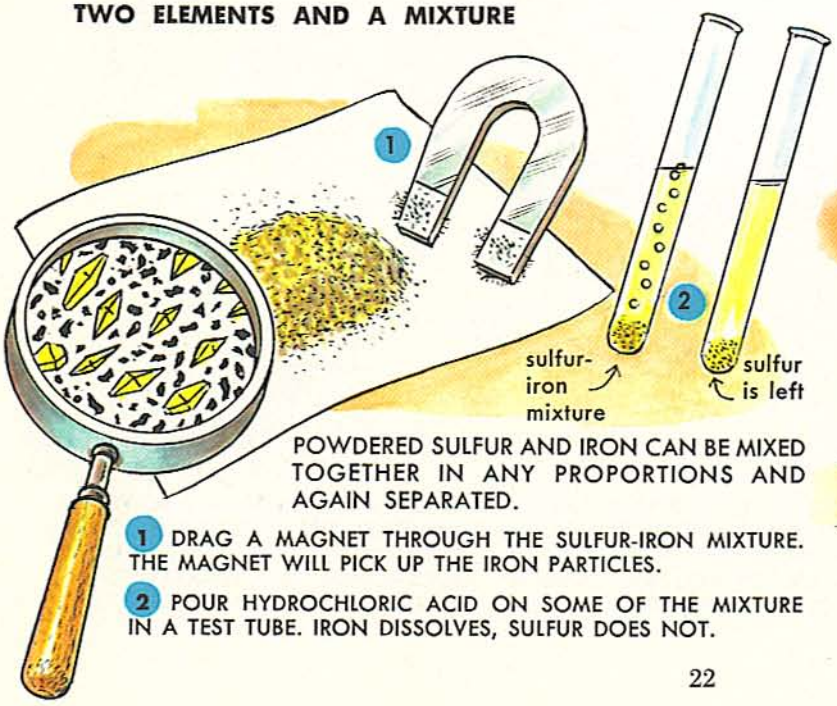
Matter has three distinct forms. Iron, for instance, is a SOLID. Water is a LIQUID. Air has the form of a GAS.

If you should take iron and divide it again and again until you couldn't divide it any further, every tiny particle would still be iron. A thing that consists of one kind of matter only is called an ELEMENT.

Take water, on the other hand. You will learn to break water up into two kinds of matter — each of them an element. A thing in which two or more elements are combined chemically is called a COMPOUND. In a compound the proportions of the different elements that make it up are always exactly the same.

Air also consists of different kinds of matter, but they are not combined chemically — they are simply mixed together. When you make a MIXTURE, you can mix the ingredients together in any proportions that suit you.

TWO ELEMENTS AND A MIXTURE



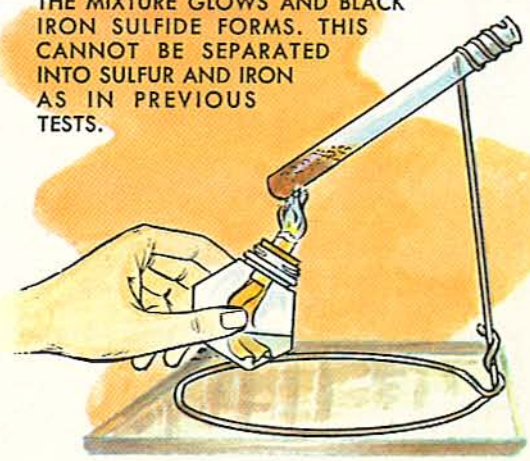
POWDERED SULFUR AND IRON CAN BE MIXED TOGETHER IN ANY PROPORTIONS AND AGAIN SEPARATED.

1 DRAG A MAGNET THROUGH THE SULFUR-IRON MIXTURE. THE MAGNET WILL PICK UP THE IRON PARTICLES.

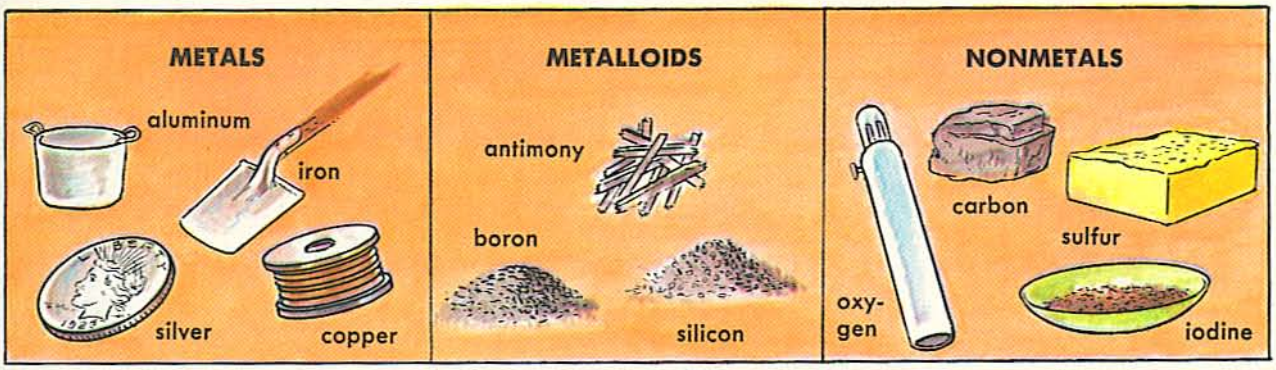
2 POUR HYDROCHLORIC ACID ON SOME OF THE MIXTURE IN A TEST TUBE. IRON DISSOLVES, SULFUR DOES NOT.

MAKING A COMPOUND

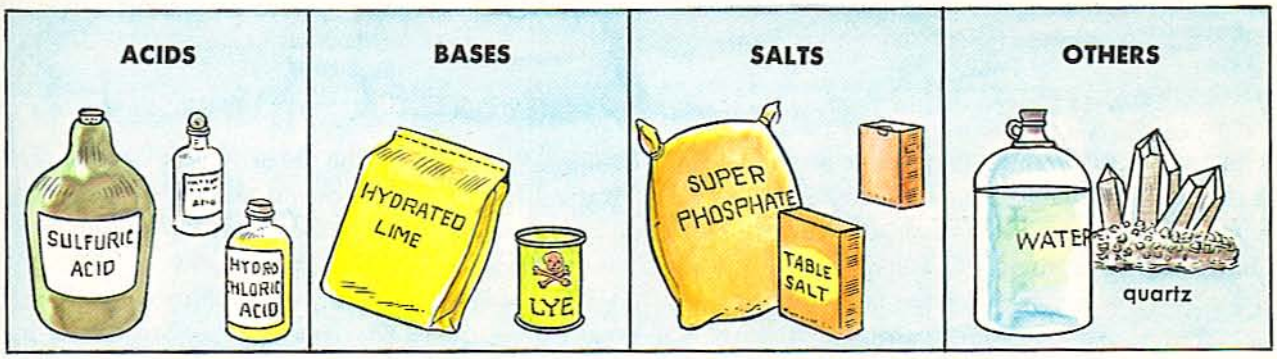
MIX TOGETHER 2 g OF FLOWERS OF SULFUR AND 3.5 g OF IRON FILINGS. PLACE MIXTURE IN A DAMAGED TEST TUBE. HEAT. SHORTLY A CHEMICAL REACTION TAKES PLACE. THE MIXTURE GLOWS AND BLACK IRON SULFIDE FORMS. THIS CANNOT BE SEPARATED INTO SULFUR AND IRON AS IN PREVIOUS TESTS.



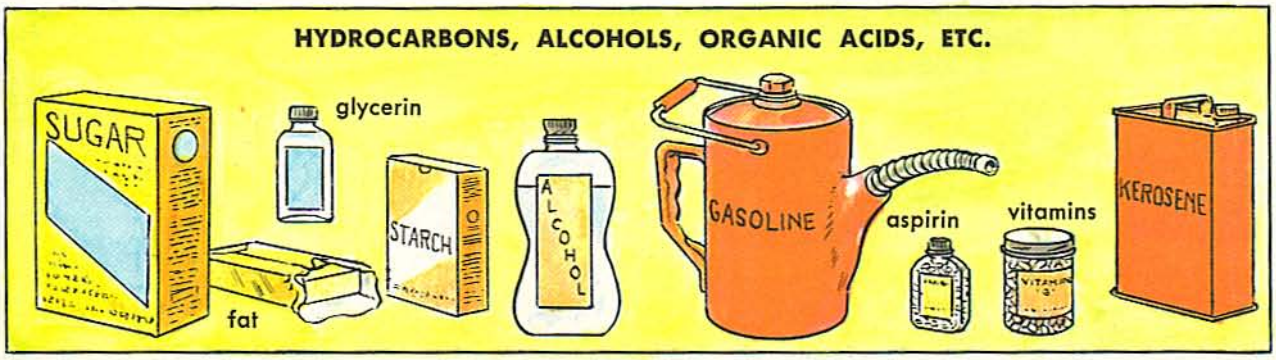
ELEMENTS ARE SUBSTANCES THAT CONSIST OF ONE KIND OF MATTER ONLY. THEY CAN BE DIVIDED INTO METALS, METALLOIDS (METAL-LIKE), NONMETALS. SEVERAL OF THE NONMETALS ARE GASES.



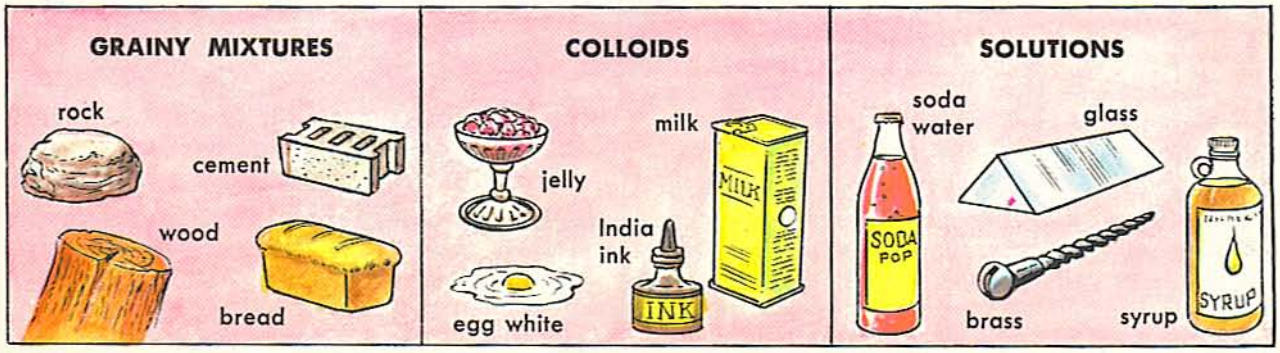
COMPOUNDS—INORGANIC. ALL COMPOUNDS CONSIST OF TWO OR MORE ELEMENTS. INORGANIC COMPOUNDS (WITH A FEW EXCEPTIONS) ARE THOSE THAT DO NOT CONTAIN THE ELEMENT CARBON.

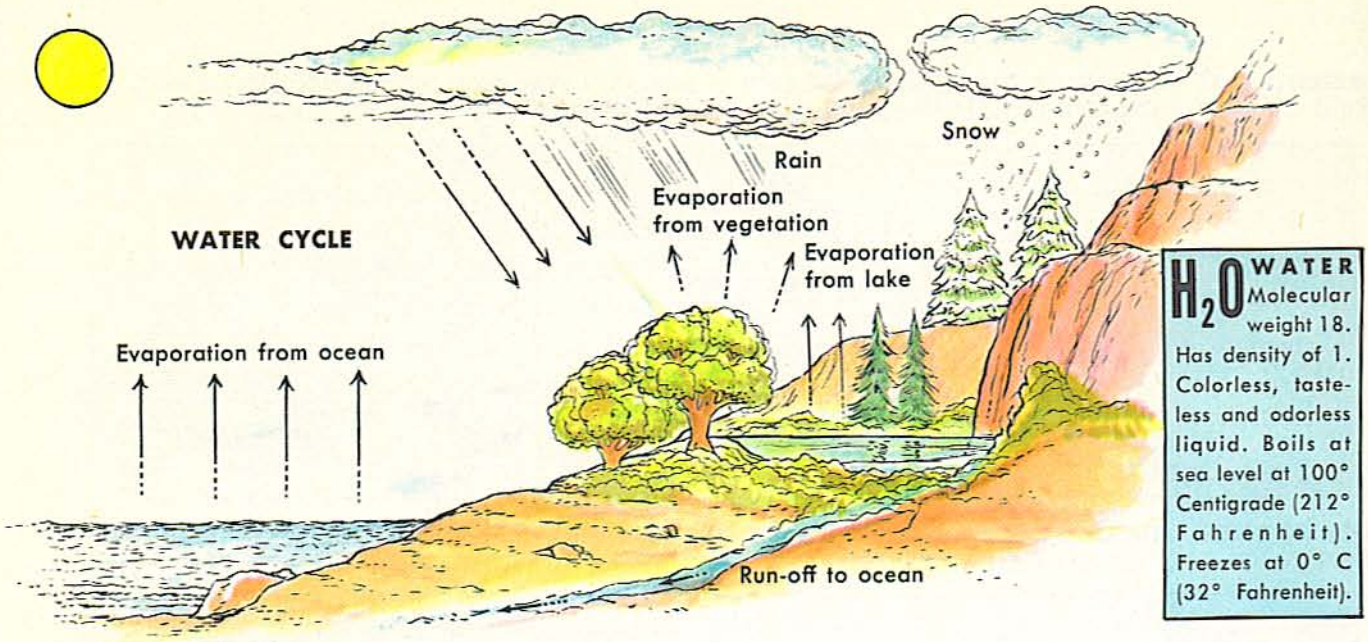


CARBON COMPOUNDS—ORGANIC. ORIGINALLY, COMPOUNDS MADE BY LIVING THINGS (PLANTS AND ANIMALS) WERE CALLED "ORGANIC." TODAY ORGANIC CHEMISTRY COVERS THE CARBON COMPOUNDS.



MIXTURES CAN CONSIST OF ELEMENTS OR COMPOUNDS. SOME MIXTURES ARE COARSE. SOME (COLLOIDS) CONTAIN TINY PARTICLES. STILL OTHERS (SOLUTIONS) ARE OF SAME STRUCTURE THROUGHOUT.





Water—Our Most Important Compound

YES, WATER is the most important of all chemical compounds. Without it, there would be no life — all human beings and all animals would thirst to death, and all plant life would wilt and die.

Fortunately, water is also the most common compound in the world. Almost three quarters of the earth's surface is covered by water. This water is forever traveling. It is turned into invisible vapor by evaporation from oceans and lakes and growing things. When cooled, the vapor forms clouds of tiny water drops. Further cooling makes the drops fall to earth as rain or snow that fill up rivers and lakes and oceans and continue the water cycle.

Chemists use nature's method to produce chemically pure water. They turn ordinary tap water into steam by boiling, then turn the steam back into water by cooling. This process is called distillation and the water is called distilled water.

WATER AS A SOLVENT

THE MOST IMPORTANT FUNCTION OF WATER IN CHEMICAL EXPERIMENTS IS AS A SOLVENT—THAT IS, A LIQUID IN WHICH CHEMICALS MAY BE DISSOLVED. FIND OUT BY AN EASY EXPERIMENT WHETHER HEATING THE WATER HELPS IN DISSOLVING A CHEMICAL.

1 DROP 1 TABLESPOON WASHING SODA IN ½ GLASS OF COLD WATER. STIR. PART OF THE SODA DISSOLVES SLOWLY.

2 REPEAT WITH HOT WATER. SODA DISSOLVES QUICKLY. HOT WATER IS USUALLY FASTER THAN COLD FOR PREPARING A SOLUTION.

WATER AS A CATALYST

WATER HELPS BRING ABOUT MANY CHEMICAL REACTIONS WITHOUT ITSELF ENTERING INTO THEM. A SUBSTANCE THAT ACTS THIS WAY IS CALLED A CATALYST.

1 PLACE 1 TEASPOON DRY BAKING POWDER IN SMALL JAR. ATTACH WIRE TO CANDLE. LIGHT CANDLE AND LOWER IT INTO JAR. CANDLE GOES ON BURNING.

2 NOW POUR WARM WATER ON THE BAKING POWDER. A CHEMICAL REACTION MAKES THE POWDER FOAM. THE GAS RELEASED IS CARBON DIOXIDE. IT MAKES CANDLE FLAME FLICKER AND GO OUT.

ELECTROLYSIS OF WATER

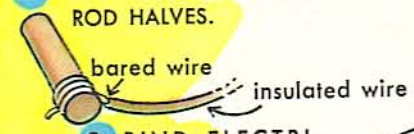
ELECTRICITY CAN BE USED TO BREAK WATER APART INTO THE TWO ELEMENTS OF WHICH IT CONSISTS —THE GASES HYDROGEN AND OXYGEN.

YOU CAN GET THE REQUIRED ELECTRICITY FROM THREE OR FOUR ORDINARY FLASHLIGHT BATTERIES. YOU WILL ALSO NEED TWO PIECES OF INSULATED COPPER WIRE AND TWO "ELECTRODES" MADE FROM CARBON RODS.

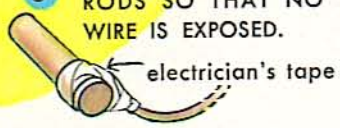
Making Electrodes

1 SCORE THE MIDDLE OF THE CARBON ROD FROM AN OLD FLASHLIGHT BATTERY, USING A FILE. BREAK THE ROD INTO TWO PIECES.

2 BARE THE WIRE FOR 2" AT EACH END OF TWO 18" LENGTHS OF INSULATED WIRE. TIE ONE BARED WIRE AROUND END OF EACH OF CARBON ROD HALVES.



3 BIND ELECTRICIAN'S TAPE FIRMLY AROUND CARBON RODS SO THAT NO WIRE IS EXPOSED.



MATERIALS FOR EXPERIMENTS

AN ORDINARY FLASHLIGHT BATTERY WILL GIVE YOU MATERIALS YOU NEED FOR EXPERIMENTS ON THIS AND SEVERAL FOLLOWING PAGES.

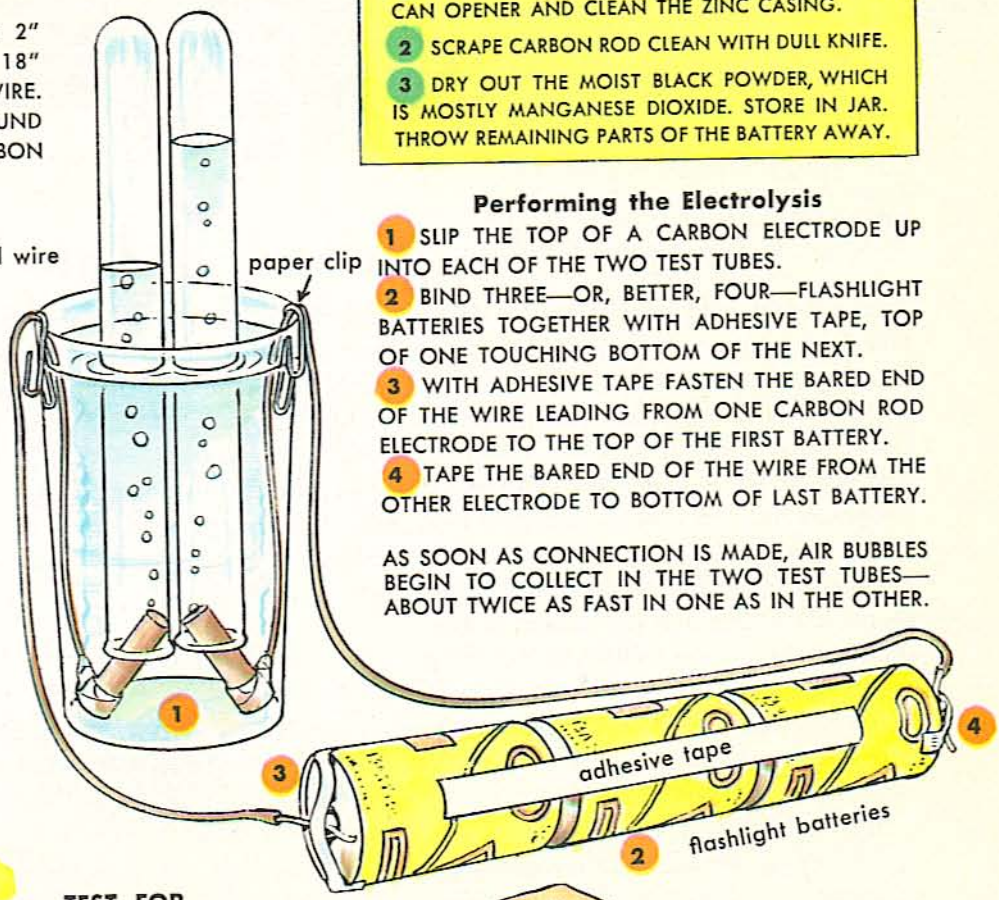
1 OPEN UP BATTERY CASE CAREFULLY WITH A CAN OPENER AND CLEAN THE ZINC CASING.

2 SCRAPE CARBON ROD CLEAN WITH DULL KNIFE.

3 DRY OUT THE MOIST BLACK POWDER, WHICH IS MOSTLY MANGANESE DIOXIDE. STORE IN JAR. THROW REMAINING PARTS OF THE BATTERY AWAY.

Setting up Electrolysis

WATER IS A POOR CONDUCTOR OF ELECTRICITY—SO YOU DISSOLVE 1 TABLESPOON OF WASHING SODA IN 1 PINT OF WATER AND FILL A WATER GLASS AND TWO TEST TUBES WITH THIS SOLUTION. THEN SET UP THE APPARATUS AS SHOWN AT RIGHT.



Performing the Electrolysis

1 SLIP THE TOP OF A CARBON ELECTRODE UP INTO EACH OF THE TWO TEST TUBES.

2 BIND THREE—OR, BETTER, FOUR—FLASHLIGHT BATTERIES TOGETHER WITH ADHESIVE TAPE, TOP OF ONE TOUCHING BOTTOM OF THE NEXT.

3 WITH ADHESIVE TAPE FASTEN THE BARED END OF THE WIRE LEADING FROM ONE CARBON ROD ELECTRODE TO THE TOP OF THE FIRST BATTERY.

4 TAPE THE BARED END OF THE WIRE FROM THE OTHER ELECTRODE TO BOTTOM OF LAST BATTERY.

AS SOON AS CONNECTION IS MADE, AIR BUBBLES BEGIN TO COLLECT IN THE TWO TEST TUBES—ABOUT TWICE AS FAST IN ONE AS IN THE OTHER.

TEST FOR HYDROGEN

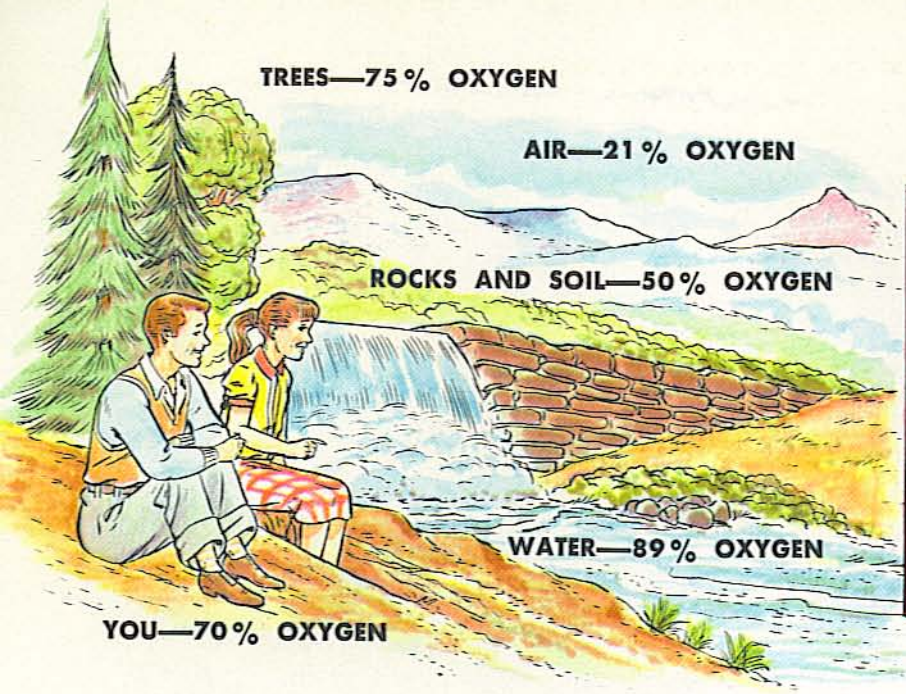
1 WITH YOUR THUMB, CLOSE THE MOUTH OF THE TEST TUBE FIRST FILLED WITH GAS. LIFT THE TUBE OUT OF THE WATER, MOUTH DOWN.

2 BRING LIGHTED MATCH TO THE MOUTH OF THE TUBE. CONTENTS BURN WITH A SOFT "POP!" THIS IS THE TEST FOR HYDROGEN.

TEST FOR OXYGEN

1 WHEN SECOND TUBE IS FULL OF GAS, CLOSE ITS MOUTH WITH YOUR THUMB. LIFT THE TUBE OUT OF THE WATER WITH MOUTH UP.

2 LIGHT A BROOMSTRAW. BLOW OUT THE FLAME. BRING THE GLOWING END DOWN IN THE TEST TUBE. GLOWING EMBER BURSTS INTO BRIGHT FLAME. THIS IS TEST FOR OXYGEN.



O OXYGEN Element 8. Atomic wt. 16. Colorless, odorless gas, supporting combustion (burning). 1.1 weight of air. Slightly soluble in water—3 volumes in 100 volumes at 20° C. Oxygen is the most common element on earth.

Oxygen—The Breath of Life

If you could hold your breath for a few minutes so that no air could get into your lungs, you would die.

For thousands of years, people have known that no human being can live without air. But it was not until Karl Scheele, a Swedish chemist, in 1772, and Joseph Priestley, an Englishman, in 1774, discovered and described oxygen that people knew that it is the oxygen in the air that is important to life.

Both of these scientists discovered that things burn more fiercely in pure oxygen than they do in the mixture of oxygen and other gases called "air."

In the lab, oxygen is produced by driving it out of certain oxygen-containing compounds. A good one to use in the home lab is hydrogen peroxide. You can get it at a drug store in a 3% solution. Hydrogen peroxide is related to water.

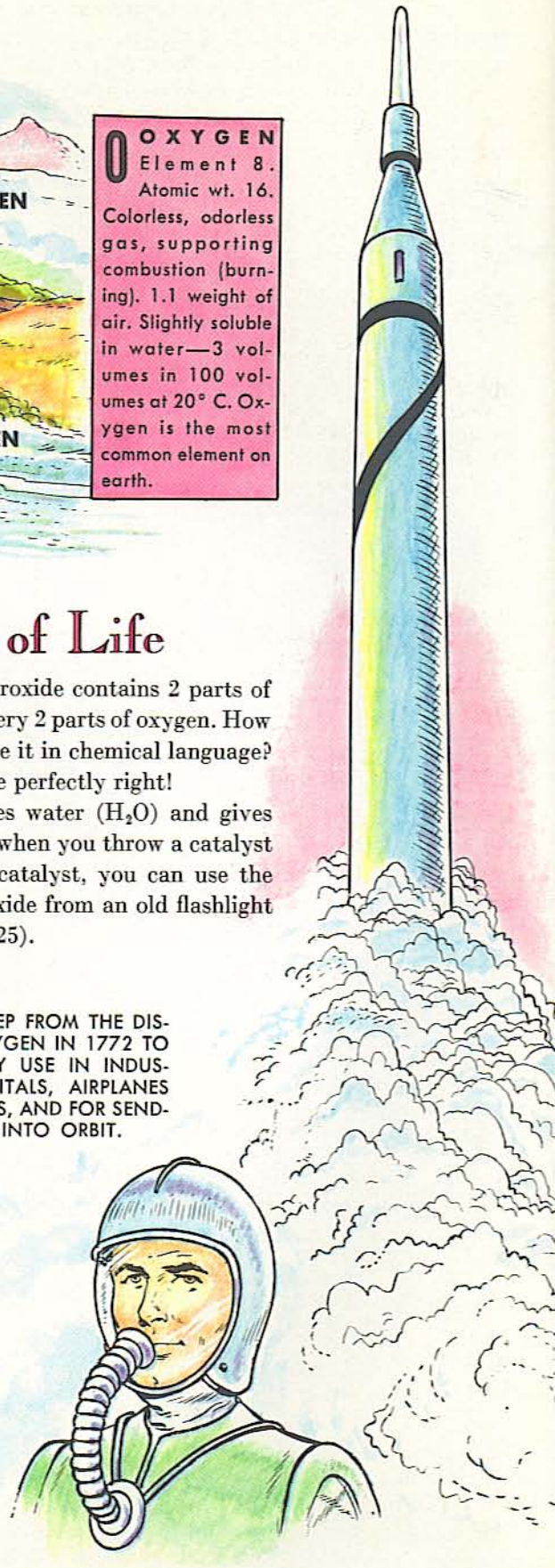
Water, as you know, consists of 2 parts of hydrogen to 1 part of oxygen. You could write it: Hydrogen 2 — Oxygen 1. That's pretty much what chemists do — except that they abbreviate the names to initials, use small numbers, and don't bother about the number 1. The formula becomes H_2O.

Hydrogen peroxide contains 2 parts of hydrogen to every 2 parts of oxygen. How would you write it in chemical language?

H_2O_2? You're perfectly right!

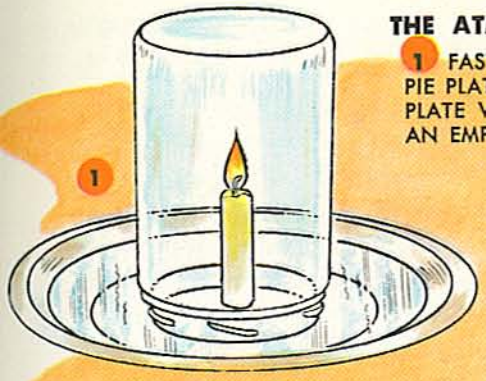
H_2O_2 becomes water (H_2O) and gives off oxygen (O) when you throw a catalyst into it. For a catalyst, you can use the manganese dioxide from an old flashlight battery (page 25).

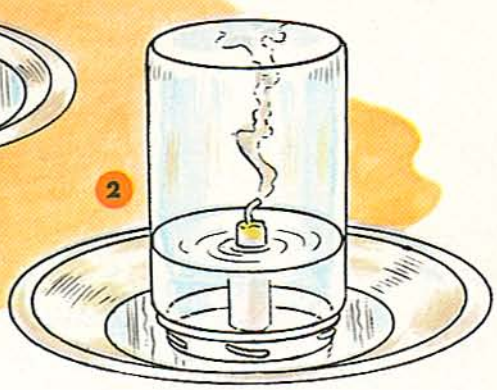
IT'S A LONG STEP FROM THE DISCOVERY OF OXYGEN IN 1772 TO ITS PRESENT-DAY USE IN INDUSTRY AND HOSPITALS, AIRPLANES AND SPACE SHIPS, AND FOR SENDING SATELLITES INTO ORBIT.



THE ATMOSPHERE CONTAINS OXYGEN

1 FASTEN A SMALL CANDLE TO MIDDLE OF PIE PLATE WITH CANDLE DRIPPINGS. FILL PIE PLATE WITH WATER. LIGHT CANDLE. PLACE AN EMPTY JAR OVER CANDLE.

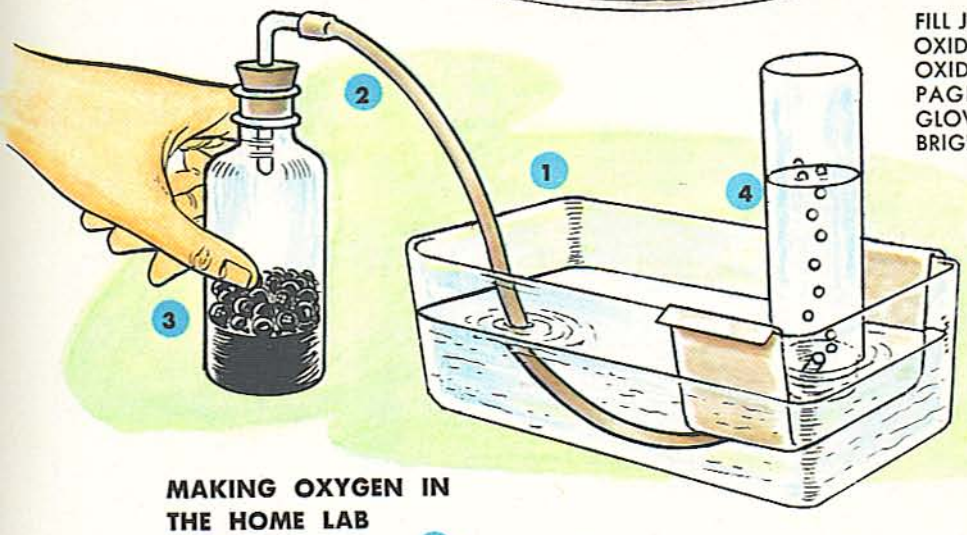


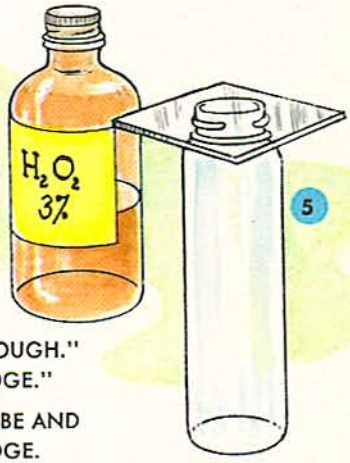


2 A MOMENT LATER, CANDLE GOES OUT. WATER RISES IN JAR TO REPLACE OXYGEN USED.

MAKING A SMALL AMOUNT OF OXYGEN

FILL JAR ¼ FULL OF 3% HYDROGEN PEROXIDE. ADD PINCH OF MANGANESE DIOXIDE FROM FLASHLIGHT BATTERY (SEE PAGE 25). TEST FOR OXYGEN WITH GLOWING BROOMSTRAW. EMBER GLOWS BRIGHTER AND MAY BURST INTO FLAME.





MAKING OXYGEN IN THE HOME LAB

1 TO COLLECT OXYGEN, YOU NEED A "PNEUMATIC TROUGH." THIS IS A DEEP, WATER-FILLED TRAY WITH METAL "BRIDGE."

2 FIT BOTTLE WITH STOPPER WITH L-SHAPED GLASS TUBE AND RUBBER TUBE LONG ENOUGH TO REACH HOLE OF BRIDGE.

3 FILL BOTTLE ¼ FULL OF 3% HYDROGEN PEROXIDE. ADD ⅛ TEASPOON OF MANGANESE DIOXIDE. PUT THE STOPPER IN.

4 FILL JAR WITH WATER AND PLACE IT UPSIDE DOWN ON THE BRIDGE IN SUCH A WAY THAT THE OXYGEN BUBBLES INTO IT AND FILLS IT BY FORCING OUT AND REPLACING THE WATER

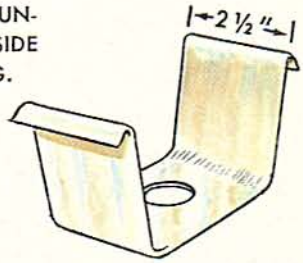
5 WHEN JAR IS FULL OF OXYGEN, SLIDE A GLASS PLATE UNDER OPENING (OR PUT STOPPER IN IT). TURN JAR RIGHT SIDE UP—QUICKLY, TO PREVENT THE OXYGEN FROM ESCAPING.

oxygen is slightly heavier than air —so keep mouth of jar up

MANY MATERIALS BURN IN OXYGEN

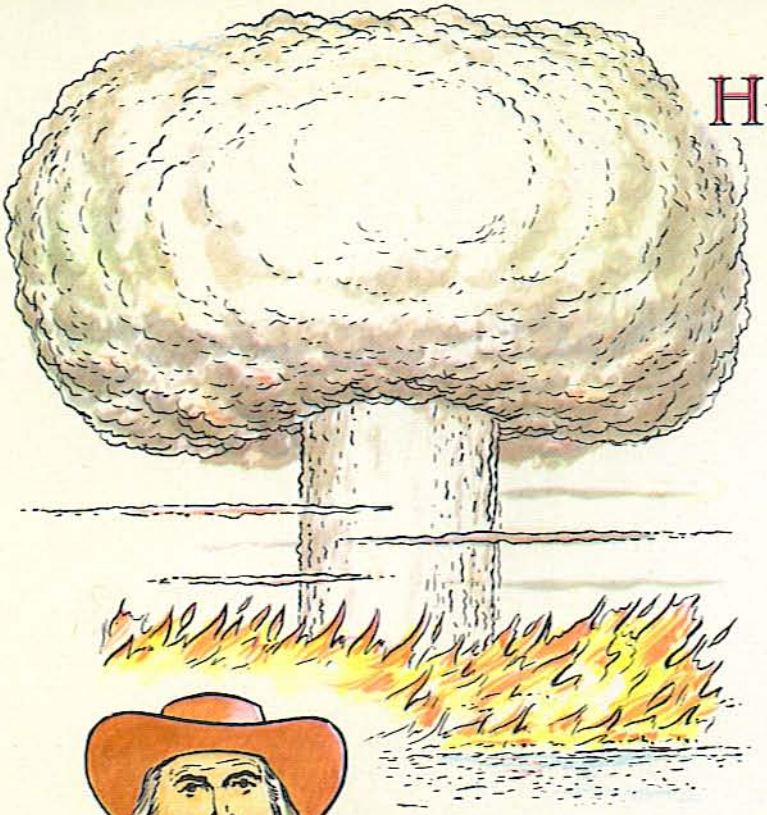
1 ATTACH TUFT OF STEEL WOOL TO WIRE. HEAT TO RED HEAT OVER ALCOHOL BURNER. LOWER INTO JAR OF OXYGEN. IRON BURSTS INTO FLAME.

2 PLACE SMALL PIECE OF SULFUR IN CROOK OF BENT STRIP OF TIN CUT FROM CAN. IGNITE SULFUR WITH MATCH. LOWER INTO JAR OF OXYGEN. SULFUR BURNS WITH A BRILLIANT, BLUE LIGHT.



"BRIDGE" FOR "PNEUMATIC TROUGH" MADE FROM 2½" STRIP OF TIN CAN.

Hydrogen—Lightest of All



HENRY CAVENDISH, WHO DISCOVERED HYDROGEN IN 1766, HAD NO IDEA OF THE ASTONISHING FORCE OF HYDROGEN WHEN RELEASED IN A BOMB.

HYDROGEN is the lightest element in existence—$\frac{1}{14}$ the weight of air. For this reason one of its early uses was for filling balloons. The first man-carrying gas balloon was sent up by the Frenchman, Jacques Charles, in 1783. The danger of using an explosive gas for this purpose was demonstrated in 1937 in the *Hindenburg* disaster, when the hydrogen-filled Zeppelin dirigible exploded on arriving at Lakehurst, New Jersey, after a trip across the Atlantic Ocean. Thirty-six people lost their lives.

Hydrogen is one of the most important of all the elements. It is found in all living things — your own body is approximately 10 per cent hydrogen. Water, as you know, is part hydrogen. So is the food you eat, the milk you drink, the clothes you wear, and such common, everyday things as gasoline and fuel oil and cooking gas.

In the home lab, you can make hydrogen by adding strips of zinc from a flashlight battery to hydrochloric acid which consists of hydrogen (H) and another gas called chlorine (Cl). The zinc forms a compound ($ZnCl_2$) with the chlorine and sets the hydrogen free (H_2).

HYDROGEN FORMS WATER WHEN IT BURNS

1. FILL TEST TUBE ¾ FULL OF HYDRO-CHLORIC ACID. ADD A COUPLE OF ZINC STRIPS. BUBBLES OF HYDROGEN FORM IMMEDIATELY.
2. CLOSE TEST TUBE WITH RUBBER STOPPER WITH GLASS TUBE DRAWN TO JET POINT. COVER APPARATUS WITH A TOWEL.
3. PLACE EMPTY TEST TUBE OVER GLASS TUBE. AFTER 1 MINUTE, TEST THIS TUBE FOR HYDROGEN WITH LIGHTED MATCH. IF TUBE "BARKS," PUT IT BACK. AFTER ANOTHER MINUTE, TRY AGAIN. WHEN SOFT "POP" TELLS YOU GAS IS PURE, LIGHT JET.
4. HOLD A COLD GLASS OVER HYDROGEN FLAME. DEW COVERING THE INSIDE OF THE GLASS SHOWS THAT WATER IS FORMING.

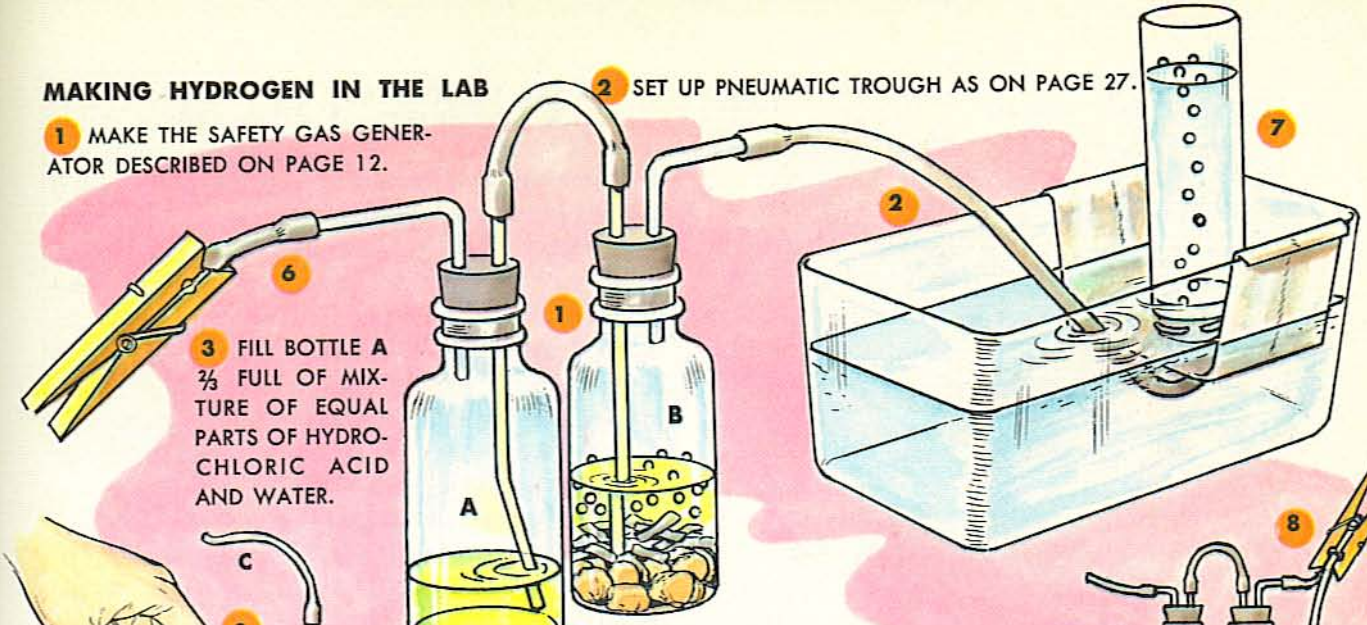
HYDROGEN IS LIGHTEST GAS KNOWN

1. FILL A POP BOTTLE ¼ FULL OF HALF-AND-HALF MIXTURE OF HYDROCHLORIC ACID AND WATER. DROP IN HALF A DOZEN ZINC STRIPS. LET NO FLAME COME NEAR!
2. FIT BALLOON ON MOUTH OF BOTTLE.
3. WHEN BALLOON IS INFLATED, TIE OPENING WITH STRING AND REMOVE FROM BOTTLE. IF PERMITTED, BALLOON WILL RISE TO CEILING INDOORS. OUTDOORS, IT WILL SOAR UP IN THE SKY.

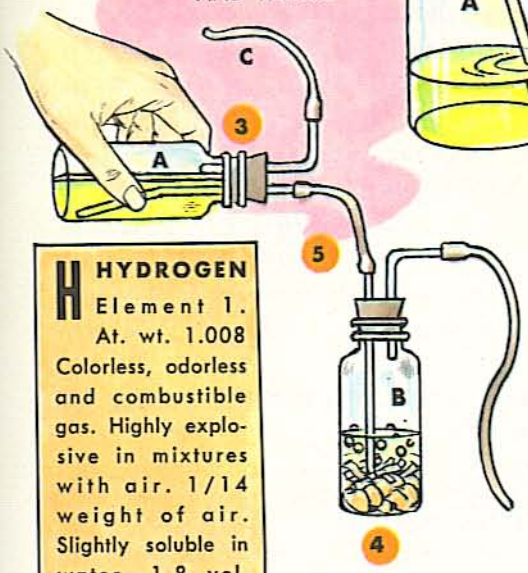
MAKING HYDROGEN IN THE LAB

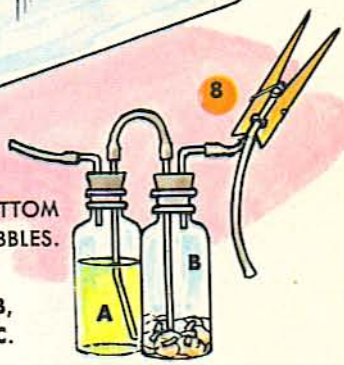
1 MAKE THE SAFETY GAS GENERATOR DESCRIBED ON PAGE 12.

2 SET UP PNEUMATIC TROUGH AS ON PAGE 27.



3 FILL BOTTLE **A** ⅔ FULL OF MIXTURE OF EQUAL PARTS OF HYDROCHLORIC ACID AND WATER.





4 PLACE 1" LAYER OF SMALL PEBBLES ON BOTTOM OF BOTTLE **B**. DROP ZINC STRIPS ON TOP OF PEBBLES. FIT STOPPERS AND TUBES AIRTIGHT.

5 POUR ACID FROM BOTTLE **A** INTO BOTTLE **B**, OR FORCE IT OVER BY BLOWING IN TUBE **C**.

6 CLOSE TUBE OF BOTTLE **A** WITH CLOTHESPIN.

7 AS SOON AS HYDROGEN STARTS TO BUBBLE UP THROUGH WATER IN PNEUMATIC TROUGH, COLLECT SOME IN WATER-FILLED TEST TUBE. WHEN FULL OF GAS, TEST IT AS DESCRIBED AT BOTTOM OF PAGE. WHEN HYDROGEN IS PURE, BUBBLE IT INTO UPSIDE-DOWN, WATER-FILLED JARS.

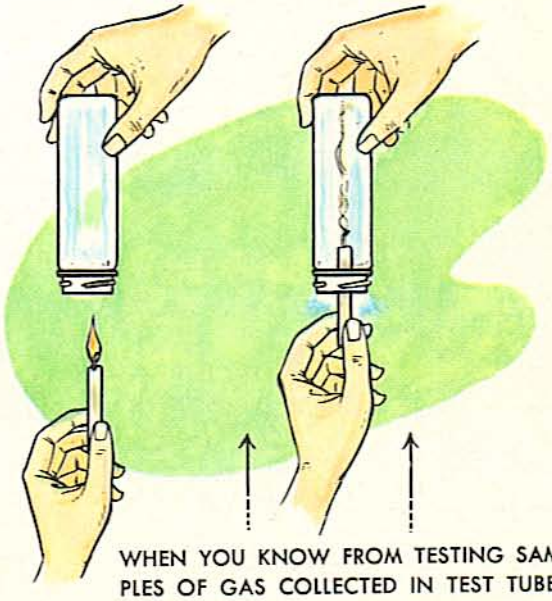
8 WHEN YOU HAVE MADE THE EXPERIMENTS YOU WANT, CHANGE SPRING CLOTHESPIN FROM RUBBER TUBE OF BOTTLE **A** TO RUBBER TUBE OF BOTTLE **B**. HYDROGEN FORCES ACID FROM BOTTLE **B** BACK INTO **A**. WHEN ACID NO LONGER TOUCHES ZINC, REACTION STOPS.

H HYDROGEN
Element 1.
At. wt. 1.008
Colorless, odorless and combustible gas. Highly explosive in mixtures with air. 1/14 weight of air. Slightly soluble in water—1.8 volumes in 100 volumes at 20° C.

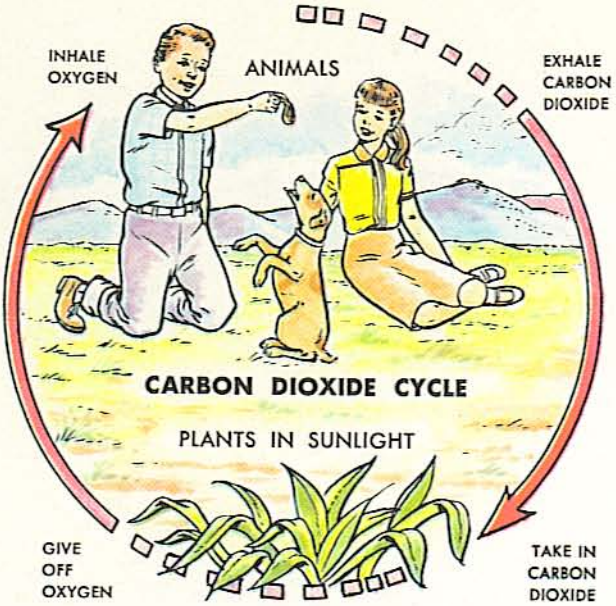
PLAYING SAFE WITH HYDROGEN

IN MIXTURES WITH AIR, HYDROGEN IS HIGHLY EXPLOSIVE. FOLLOW SAFETY RULES BELOW.

- MAKE ONLY SMALL AMOUNTS OF HYDROGEN IN THE HOME LAB. A 4-OZ. GENERATOR BOTTLE WILL GIVE YOU ALL THE HYDROGEN YOU NEED. MAKE ALL CONNECTIONS AIRTIGHT.
- TEST HYDROGEN FOR PURITY BY COLLECTING A TEST TUBE FULL OF IT AND BRINGING A LIGHTED MATCH TO MOUTH OF TUBE, AS SHOWN ON PAGE 25. HYDROGEN MIXED WITH AIR EXPLODES WITH A SHARP "BARK." PURE HYDROGEN BURNS WITH A QUIET "POP."
- KEEP FLAME AWAY FROM YOUR MAIN GENERATOR BOTTLE.
- IGNITE HYDROGEN ONLY FROM TEST TUBE GENERATOR DESCRIBED ON OPPOSITE PAGE, AND THEN ONLY AFTER YOU HAVE TESTED IT FOR PURITY.



WHEN YOU KNOW FROM TESTING SAMPLES OF GAS COLLECTED IN TEST TUBES THAT HYDROGEN IS PURE, FILL SMALL JAR WITH IT. LIFT JAR OUT OF WATER, MOUTH DOWN. BRING LIGHTED CANDLE UP INTO JAR. HYDROGEN BURNS AT MOUTH OF JAR. CANDLE GOES OUT.

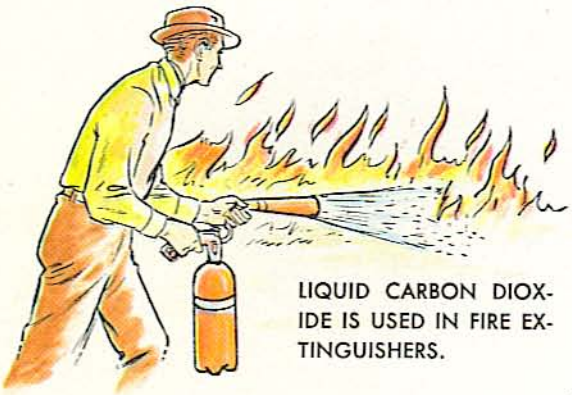


Carbon Dioxide

You have already learned in experimenting with a burning candle that when something containing carbon burns in the air, a gas, carbon dioxide (CO_2), is formed. This is one of the most important gases for human life. The reason is that green plants, in sunlight, are able to take the carbon out of the carbon dioxide in the air and, by combining it with oxygen and hydrogen from water and with various minerals in the soil, produce all the vegetable matter that humans and animals eat.

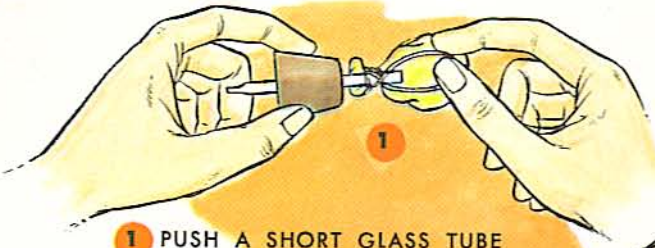
You cannot see the CO_2 in the air — but you can see it when it has been cooled and compressed into a solid block of "dry ice." When dissolved in water (H_2O), carbon dioxide (CO_2) forms a weak acid (H_2CO_3). You know the taste of this acid from soda water — the bubbles are CO_2 being set free.

Carbonic acid combines with many metals to make "carbonates." You can drive the CO_2 out of most carbonates with the help of a weak acid — even with vinegar, which is diluted acetic acid.

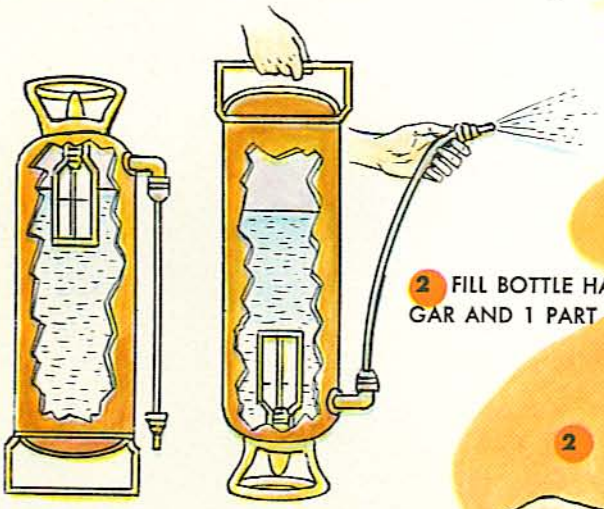


LIQUID CARBON DIOXIDE IS USED IN FIRE EXTINGUISHERS.

MAKING A FIRE EXTINGUISHER MODEL



1 PUSH A SHORT GLASS TUBE WITH A JET TIP INTO A RUBBER STOPPER. WRAP BICARBONATE OF SODA IN A SHEET OF TOILET TISSUE. ATTACH SODA PACKAGE TO TUBE WITH A RUBBER BAND.



2 FILL BOTTLE HALF FULL OF MIXTURE OF 1 PART VINEGAR AND 1 PART WATER. PUT IN THE STOPPER.

3 HOLD STOPPER FIRMLY IN PLACE WITH TWO FINGERS. TURN BOTTLE UPSIDE DOWN. THE CO_2 FORMED BY MIXING VINEGAR AND SODA DRIVES WATER OUT IN POWERFUL JET.

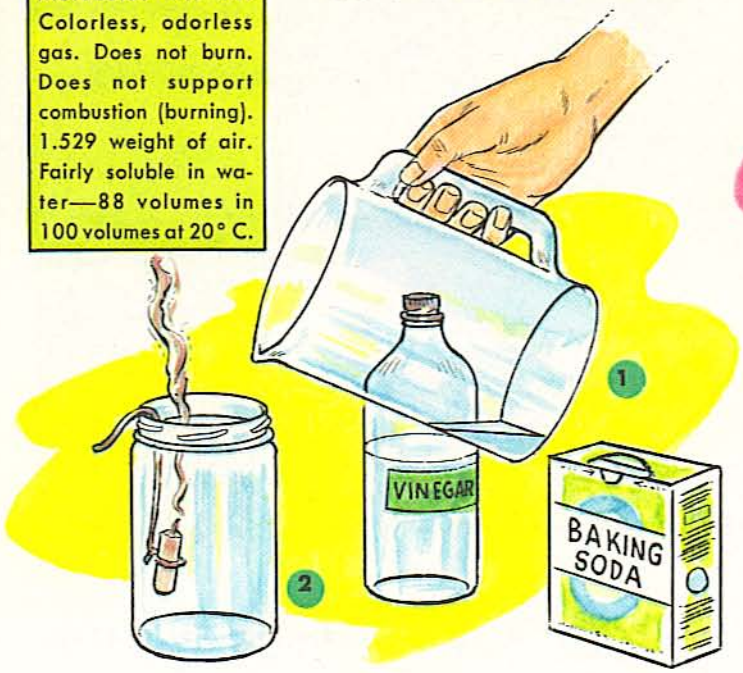
CHEMICAL FIRE EXTINGUISHERS CONTAIN SOLUTION OF BAKING SODA AND A BOTTLE OF SULFURIC ACID. WHEN TURNED UPSIDE DOWN, THE CHEMICALS MIX AND FORM CARBON DIOXIDE WHICH FORCES OUT THE WATER.

CO_2 CARBON DIOXIDE Compound.

Molecular wt. 44. Colorless, odorless gas. Does not burn. Does not support combustion (burning). 1.529 weight of air. Fairly soluble in water—88 volumes in 100 volumes at 20° C.

FEATURES OF CARBON DIOXIDE

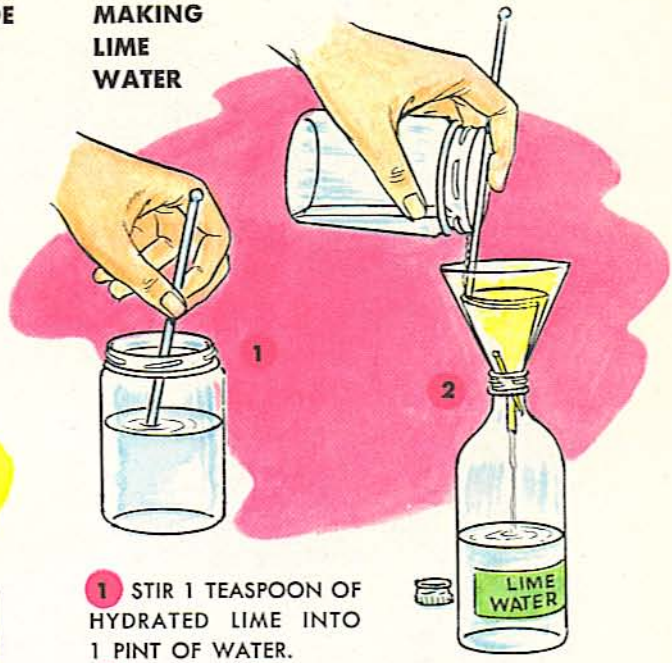
CO_2 IS HEAVIER THAN AIR AND DOES NOT SUPPORT BURNING. YOU CAN PROVE BOTH POINTS:



1 PLACE 1 TEASPOON OF BAKING SODA IN A PITCHER. POUR A SMALL AMOUNT OF WHITE VINEGAR OVER THE SODA.

2 HANG A LIGHTED CANDLE IN A JAR BY A WIRE. POUR THE CARBON DIOXIDE FORMED IN THE PITCHER INTO THE JAR THE WAY YOU WOULD POUR WATER. WHEN THE CARBON DIOXIDE REACHES THE TOP OF THE CANDLE, THE FLAME GOES OUT.

MAKING LIME WATER



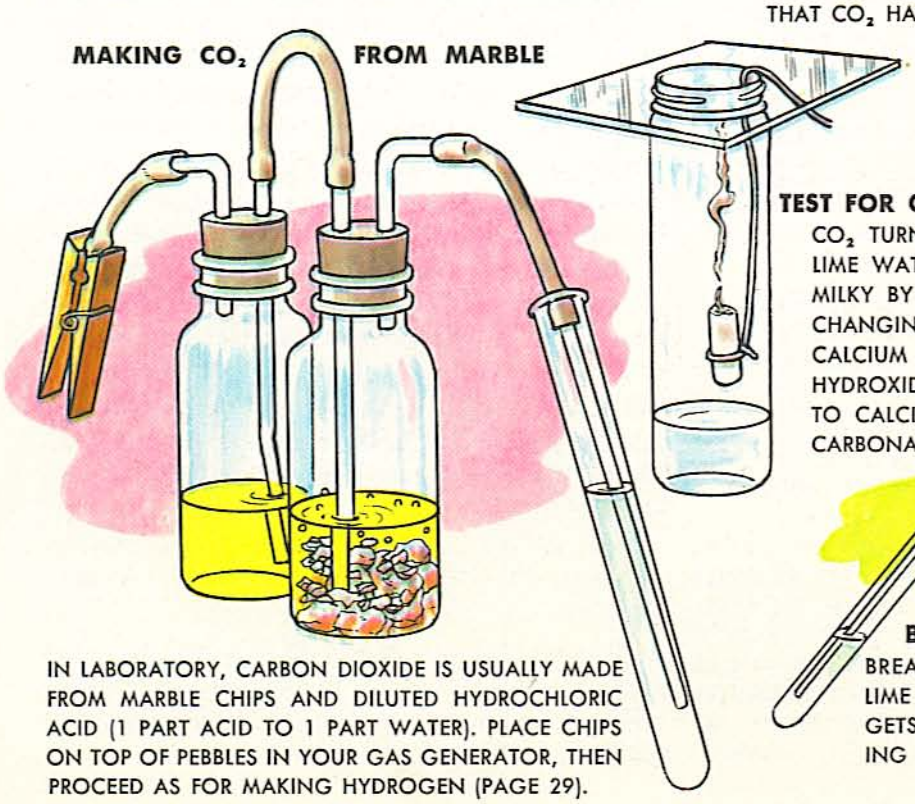
1 STIR 1 TEASPOON OF HYDRATED LIME INTO 1 PINT OF WATER.

2 LET STAND UNTIL LIME SINKS TO THE BOTTOM. FILTER LIQUID INTO A BOTTLE. CLOSE BOTTLE TIGHTLY.

BURNING PRODUCES CARBON DIOXIDE

HANG BURNING CANDLE IN JAR CONTAINING A FEW ml LIME WATER. COVER TOP WITH A GLASS PLATE. WHEN CANDLE HAS GONE OUT, SHAKE LIME WATER UP WITH THE AIR. MILKINESS PROVES THAT CO_2 HAS BEEN PRODUCED.

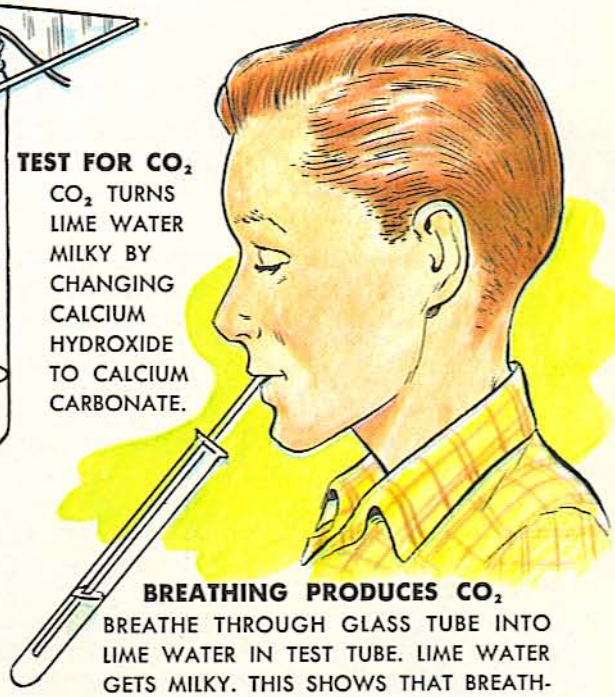
MAKING CO_2 FROM MARBLE



IN LABORATORY, CARBON DIOXIDE IS USUALLY MADE FROM MARBLE CHIPS AND DILUTED HYDROCHLORIC ACID (1 PART ACID TO 1 PART WATER). PLACE CHIPS ON TOP OF PEBBLES IN YOUR GAS GENERATOR, THEN PROCEED AS FOR MAKING HYDROGEN (PAGE 29).

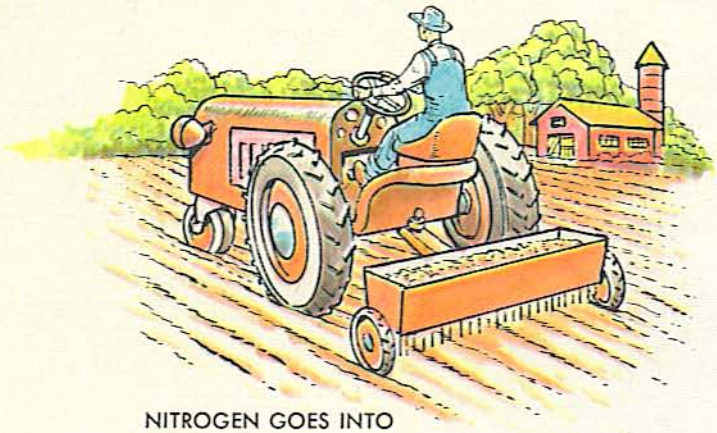
TEST FOR CO_2

CO_2 TURNS LIME WATER MILKY BY CHANGING CALCIUM HYDROXIDE TO CALCIUM CARBONATE.

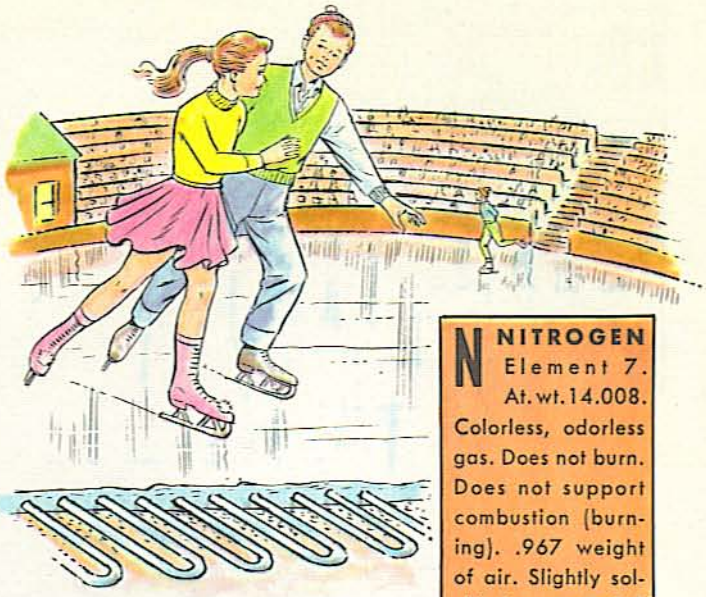


BREATHING PRODUCES CO_2

BREATHE THROUGH GLASS TUBE INTO LIME WATER IN TEST TUBE. LIME WATER GETS MILKY. THIS SHOWS THAT BREATHING IS A BURNING PROCESS.



NITROGEN GOES INTO BOTH FERTILIZERS AND EXPLOSIVES.



DRY AMMONIA GAS IS USED IN THE LARGE-SCALE PRODUCTION OF ICE.

N NITROGEN Element 7. At. wt. 14.008. Colorless, odorless gas. Does not burn. Does not support combustion (burning). .967 weight of air. Slightly soluble in water—1.5 volumes in 100 vols. at 20°C.

Nitrogen and Its Compounds

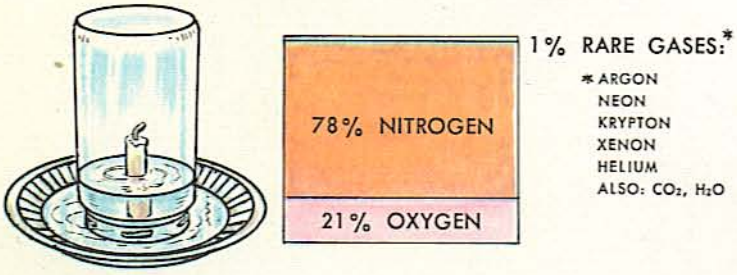
When you burn anything in the air, only about one-fifth of the air goes into chemical combination with what you are burning. The rest (except for a small fraction) does not enter into the process. It is a gas called nitrogen (N) — the most abundant free element on earth.

Nitrogen is what you might call a "lazy" element. It does not help in burning nor does it burn if you try to ignite it. It is only at high temperatures and under great pressures that a chemist can make nitrogen combine with another element, hydrogen, to form ammonia gas (NH_3), from which other nitrogen compounds can be made.

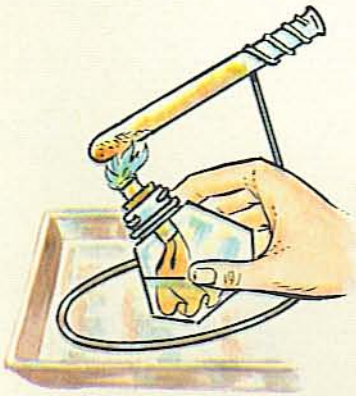
Yet, in nature, tiny bacteria on the roots of certain plants can take nitrogen from the air and make it combine with oxygen and minerals in the soil into "nitrates." And that is of tremendous importance to all of us — for all plants need nitrates if they are to thrive. If plants do not get nitrates naturally, the farmer must add them to his soil in the form of some kind of fertilizer.

You will not have much satisfaction out of working with nitrogen itself, but you will find it interesting to deal with some of its compounds — especially with ammonia gas (NH_3). You will also want to have a look at one of the half dozen combinations nitrogen makes with oxygen, the brown gas called nitrogen dioxide (NO_2).

NITROGEN FROM THE ATMOSPHERE

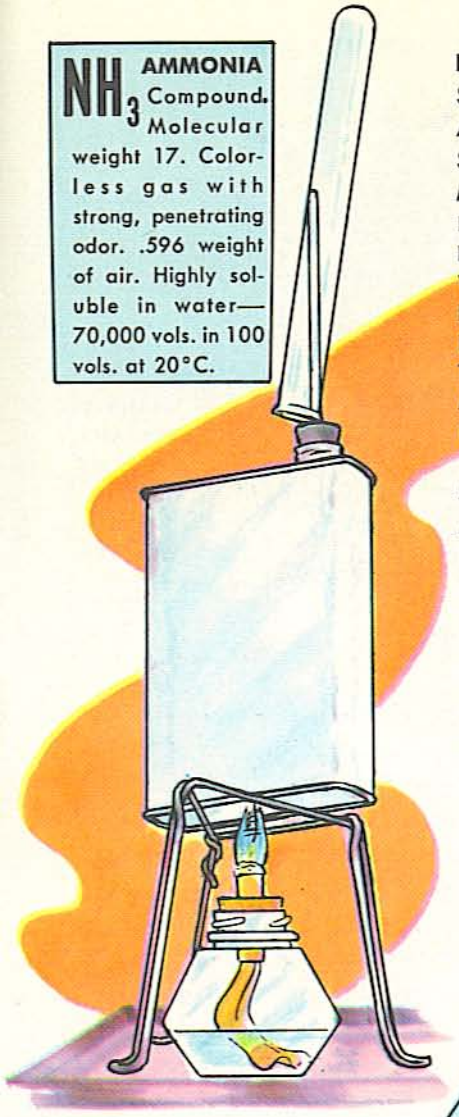


REPEAT CANDLE-BURNING EXPERIMENT ON PAGE 27. UNUSED GAS IS ALMOST ALL NITROGEN—WITH SMALL PERCENTAGE OF RARE GASES AND CARBON DIOXIDE.



NITROGEN DIOXIDE IN A WELL-VENTILATED ROOM, HEAT EQUAL AMOUNTS OF SALTPETER AND SODIUM BISULFATE IN DRY TEST TUBE. IN A MOMENT, A BROWN GAS FORMS. IT IS NITROGEN DIOXIDE. DO NOT INHALE —GAS IS VERY IRRITATING.

NH_3 AMMONIA Compound. Molecular weight 17. Colorless gas with strong, penetrating odor. .596 weight of air. Highly soluble in water—70,000 vols. in 100 vols. at 20°C.



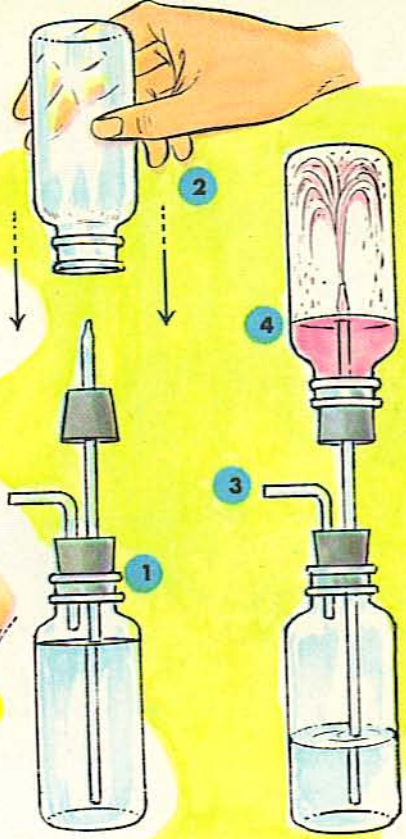
PRODUCING AMMONIA

SIMPLEST WAY OF PRODUCING AMMONIA IS TO GET IT FROM ITS SOLUTION AS HOUSEHOLD AMMONIA.

FILL PINT CAN ONE QUARTER FULL OF HOUSEHOLD AMMONIA. FIT STOPPER WITH 6" GLASS TUBE IN OPENING. PLACE TEST TUBE OVER GLASS TUBE. HEAT CAN OVER LOW FLAME. TEST TUBE IS FULL OF AMMONIA WHEN MOIST, RED LITMUS PAPER HELD AT ITS MOUTH TURNS BLUE.

SOLUBILITY OF AMMONIA

REMOVE A FILLED TEST TUBE FROM GAS GENERATOR CAN, MOUTH DOWN. CLOSE MOUTH OF TUBE WITH THUMB. OPEN TUBE UNDER WATER. AMMONIA DISSOLVES EASILY, WATER RUSHES IN AND FILLS TUBE.



THE AMMONIA FOUNTAIN

AMMONIA'S EXTRAORDINARY SOLUBILITY CAN BE SHOWN IN A SPECTACULAR DEMONSTRATION.

1 MAKE UP APPARATUS AS SHOWN IN ILLUSTRATION. FILL IT WITH WATER. ADD 5 DROPS OF PHENOLPHTHALEIN SOLUTION.

2 FILL DRY, EMPTY BOTTLE WITH AMMONIA FROM GENERATOR CAN. KEEPING BOTTLE UPSIDE DOWN, PLACE IT FIRMLY ON TOP STOPPER OF APPARATUS.

3 BLOW INTO L-SHAPED GLASS TUBE TO DRIVE A FEW DROPS OF WATER UP INTO THE UPPER BOTTLE.

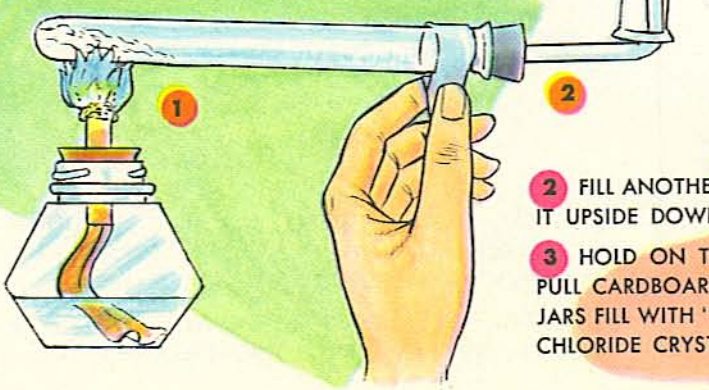
4 SUDDENLY, WATER SPURTS FROM LOWER BOTTLE UP INTO UPPER BOTTLE IN A FOUNTAIN THAT TURNS PINK AS AMMONIA REACTS ON PHENOLPHTHALEIN.

MOIST, RED LITMUS PAPER TURNS BLUE IN AMMONIA.

MAKING AMMONIA FROM SAL AMMONIAC

1 ON A PIECE OF PAPER, MIX 1 PART OF SAL AMMONIAC WITH 2 PARTS OF HYDRATED LIME. ADD A FEW DROPS OF WATER. DROP MIXTURE INTO A TEST TUBE. PROVIDE TUBE WITH STOPPER AND L-SHAPED GLASS TUBE. THEN HEAT OVER LOW FLAME.

2 COLLECT AMMONIA IN DRY TEST TUBE. TEST IT WITH LITMUS PAPER AND FOR SOLUBILITY.

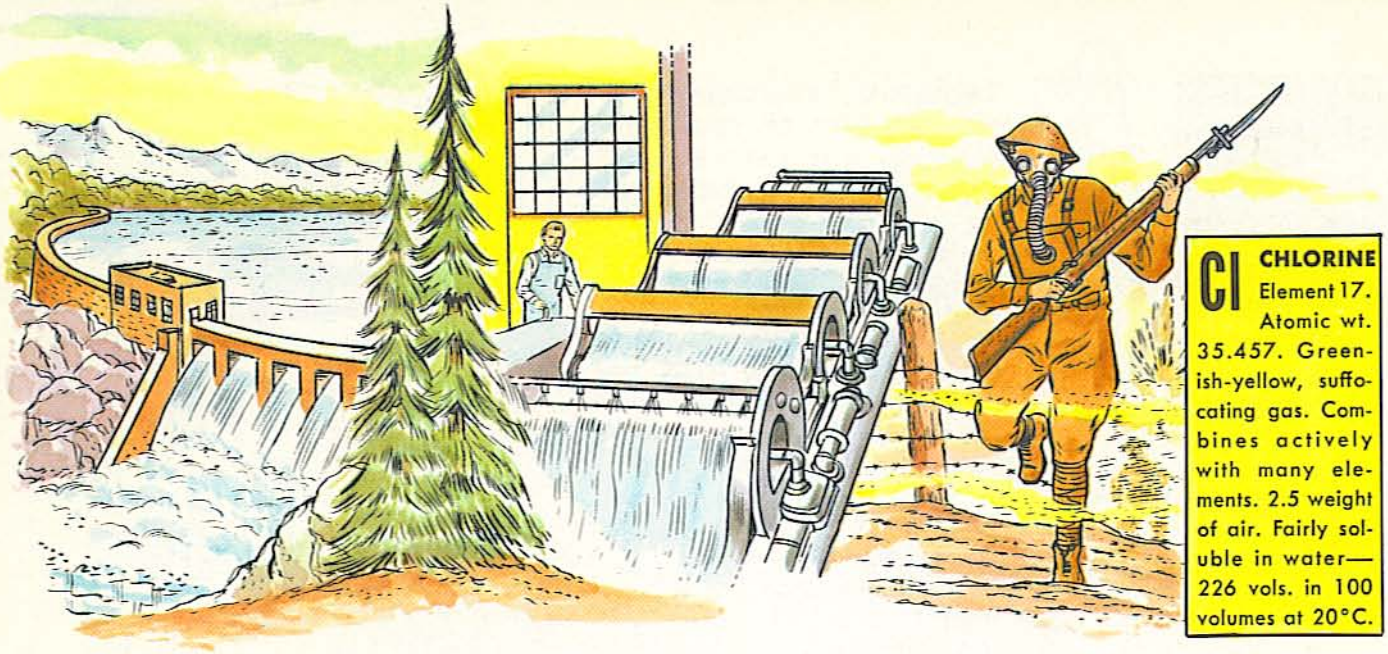


THE WHITE SMOKE MYSTERY

1 MOISTEN INSIDE OF JAR WITH SMALL AMOUNT OF HYDROCHLORIC ACID. POUR EXCESS ACID BACK INTO ITS BOTTLE. COVER JAR WITH SQUARE OF CARDBOARD.

2 FILL ANOTHER JAR WITH AMMONIA. PLACE IT UPSIDE DOWN ON CARDBOARD.

3 HOLD ON TO AMMONIA-FILLED JAR AND PULL CARDBOARD AWAY. IMMEDIATELY, BOTH JARS FILL WITH "SMOKE" OF TINY AMMONIUM CHLORIDE CRYSTALS.



Cl **CHLORINE** Element 17. Atomic wt. 35.457. Greenish-yellow, suffocating gas. Combines actively with many elements. 2.5 weight of air. Fairly soluble in water—226 vols. in 100 volumes at 20°C.

Chlorine—Friend and Foe

CHLORINE IS a gas of great importance. We wouldn't be certain of safe drinking water in our cities if it weren't for chlorine — a small amount of it in the water kills the dangerous germs that may lurk in it. Chlorine is also used extensively in bleaching.

Chlorine is a friendly gas when it is used correctly. But it is dangerous when used improperly because it affects the lungs. As a "poison gas" it caused many casualties in World War I.

You can produce chlorine as a greenish-yellow gas by driving it out of one of its compounds — hydrochloric acid (HCl), which consists of hydrogen (H) and chlorine (Cl), or a common laundry bleach ("Clorox" or others), which is a solution of sodium hypochlorite (NaClO).

Have a bottle of diluted household ammonia (90% water, 10% household ammonia) on hand. Sniff this if you get too strong a whiff of chlorine.

NOTE: Perform these experiments out-of-doors or before an open window. Be careful not to breathe fumes.

MAKING TEST PAPER FOR CHLORINE

MIX .5 g (1/4 TEASPOON) STARCH WITH 30 ml WATER. BRING TO BOIL. DISSOLVE IN MIXTURE A SMALL AMOUNT OF POTASSIUM IODIDE (AS MUCH AS TWO GRAINS OF RICE). DIP STRIPS OF FILTER PAPER IN MIXTURE; THEN DRY THEM.

CHLORINE FROM HYDROCHLORIC ACID

1 Put .5 g (1/8 TEASPOON) MANGANESE DIOXIDE INTO TEST TUBE. ADD 3 ml (1/8 TEST TUBE) UNDILUTED HYDROCHLORIC ACID. HEAT GENTLY. CHLORINE FORMS. WAFT A LITTLE CAREFULLY TOWARD YOU FOR A SNIFF.

2 TEST GAS BY HOLDING MOISTENED STARCH-IODIDE PAPER AT MOUTH OF TUBE. PAPER TURNS BLUE.

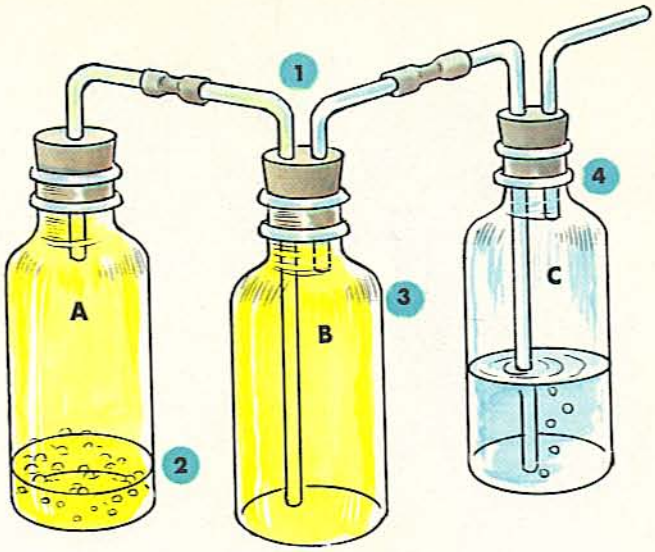
MAKING CHLORINE IN THE HOME LAB

1 MAKE APPARATUS SHOWN AT RIGHT. POUR 1 INCH OF LIQUID BLEACH (CLOROX) INTO BOTTLE **A**. BOTTLE **B** IS EMPTY. BOTTLE **C** HAS WATER IN WHICH ½ TEASPOON LYE IS DISSOLVED.

2 TAKE STOPPER OUT OF BOTTLE **A**. DROP IN ½ TEASPOON SODIUM BISULFATE (SANI-FLUSH). REPLACE THE STOPPER.

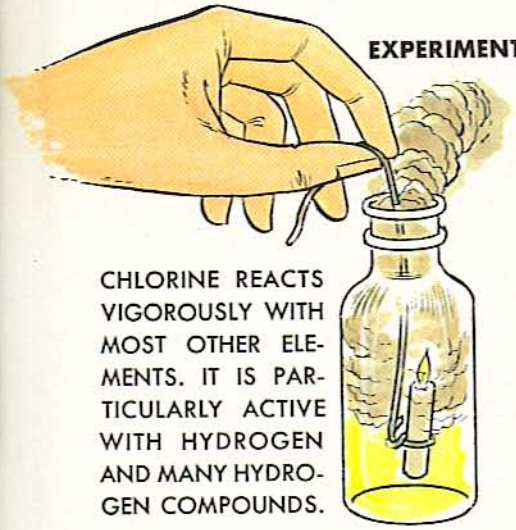
3 CHLORINE GAS FORMS AND FILLS **B**.

4 LYE WATER IN BOTTLE **C** ABSORBS EXCESS OF CHLORINE GAS.



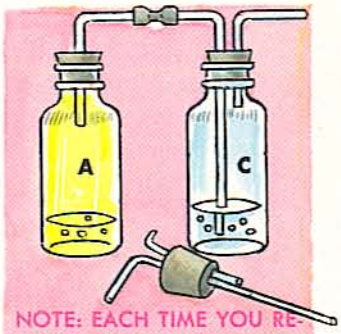
WHEN REACTION SLOWS ADD MORE SODIUM BISULFATE

EXPERIMENTS WITH CHLORINE



CHLORINE REACTS VIGOROUSLY WITH MOST OTHER ELEMENTS. IT IS PARTICULARLY ACTIVE WITH HYDROGEN AND MANY HYDROGEN COMPOUNDS.

LOWER A BURNING CANDLE INTO A BOTTLE OF CHLORINE GAS. A DENSE SMOKE OF CARBON IS FORMED. THE CHLORINE COMBINES WITH THE HYDROGEN OF THE CANDLE AND SETS THE CARBON IN IT FREE AS SOOT.



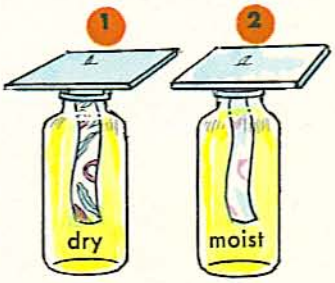
NOTE: EACH TIME YOU REMOVE THE GAS-COLLECTING BOTTLE **B** FOR EXPERIMENT, CONNECT BOTTLES **A** AND **C** TO PREVENT CHLORINE FROM GETTING OUT IN THE ROOM.

TO SHOW THE SOLUBILITY OF CHLORINE, POUR A SMALL AMOUNT OF WATER INTO A CHLORINE-FILLED BOTTLE. CLOSE THE BOTTLE MOUTH WITH YOUR PALM. SHAKE. THE CHLORINE DISSOLVES AND THE BOTTLE STICKS TO YOUR PALM FROM THE SUCTION CREATED.

CHLORINE WILL COMBINE DIRECTLY WITH SEVERAL METALS. IRON ACTUALLY BURNS IN CHLORINE GAS!

FASTEN A SMALL WAD OF STEEL WOOL TO A PIECE OF WIRE. HEAT IT WITH A MATCH AND LOWER IT INTO CHLORINE-FILLED BOTTLE. A HEAVY BROWN SMOKE OF IRON CHLORIDE POURS OUT.

CHLORINE HAS GREAT USE IN BLEACHING COTTON AND LINEN AND WOOD PULP. YET IT IS NOT THE CHLORINE THAT PERFORMS THE BLEACHING.



1 FILL A BOTTLE WITH CHLORINE GAS. HANG IN IT (FROM A CORK OR FROM A PIECE OF CARDBOARD) A STRIP OF DRY, BRIGHTLY COLORED COTTON CLOTH. NOTHING HAPPENS. COLOR OF CLOTH IS NOT AFFECTED.

2 MOISTEN THE CLOTH AND AGAIN HANG IT IN THE CHLORINE. SOON THE COLORS FADE—ONLY TRULY "FAST" COLORS REMAIN. CHLORINE, IN CONTACT WITH WATER, COMBINES WITH THE HYDROGEN AND LIBERATES OXYGEN. THE LIBERATED OXYGEN DOES THE BLEACHING.

Chemical Shorthand

So FAR YOU have experimented with oxygen and hydrogen, carbon dioxide and nitrogen, and chlorine; you have also separated water into the two elements of which it consists, and have combined the two elements iron and sulfur into a chemical compound. In taking notes of your experiments you are certain to have learned that it is much quicker to write "H" than "hydrogen," and easier to write "CO_2" than "carbon dioxide." Before long, it will seem the simplest and most logical thing in the world to use these abbreviations of the names of the different elements rather than the full names.

Yet it took chemists hundreds of years before they settled on this uniform method of writing out their chemical formulas.

In the early days of chemistry no one bothered to do much writing about it. But it became necessary for the alchemists to write down their experiments — how else could they retrace their steps in case they actually hit upon the gold they were seeking? They invented a whole line of complicated symbols that only they could understand.

As chemists delved deeper and deeper into the mysteries of matter it became more and more important for them to write out their experiments in such a way that all other chemists would know what they were trying to explain.

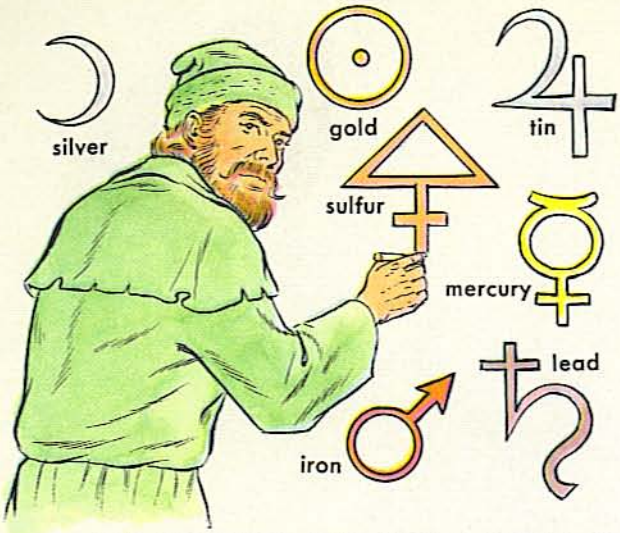
The first to invent a usable system was John Dalton, an English scientist. The invention was almost forced upon him.

In his study of chemistry he had become convinced that all chemical reactions could be explained in terms of the tiniest possible part of one element reacting with the tiniest possible part of another. These particles he called "atoms." The smallest possible part of the compound that resulted he called a "compound atom"— today we call it a "molecule."

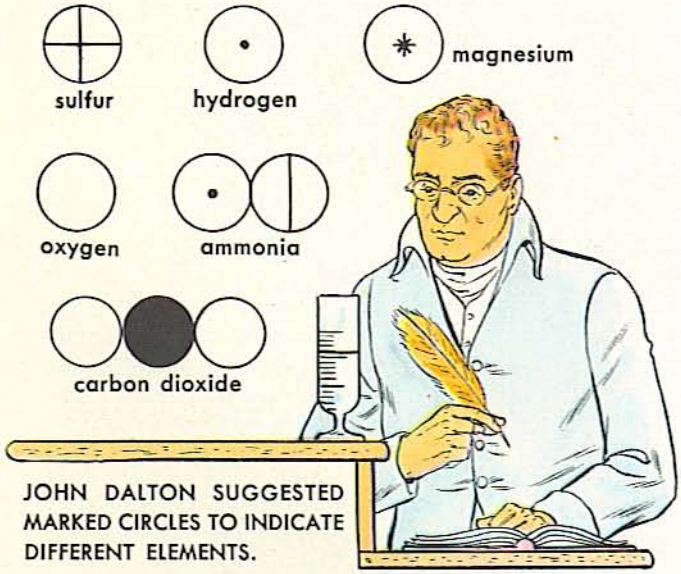
To explain his "atomic theory" Dalton made use of circles, each with a marking to indicate a specific element. These circles served to explain Dalton's theory but they were too difficult to work with to show complicated chemical reactions.

A Swedish chemist, Jons Jakob Berzelius, worked out a simpler system — the same system scientists use today.

For his symbols he took the first letter of the Latin name of each element — C for "carbo," S for "sulfur." Where two names started with the same letter, he added a small letter to one of the symbols to



THE ALCHEMISTS USED FANCIFUL FIGURES TO REPRESENT THE CHEMICALS WITH WHICH THEY WORKED.



JOHN DALTON SUGGESTED MARKED CIRCLES TO INDICATE DIFFERENT ELEMENTS.

JONS JAKOB BERZELIUS DEVELOPED THE SYSTEM USED TODAY, IN WHICH THE NAMES OF ELEMENTS ARE ABBREVIATED.

distinguish the two elements from each other — he used Ca for "calcium," for instance, to distinguish it from carbon (C).

But Berzelius went an important step further.

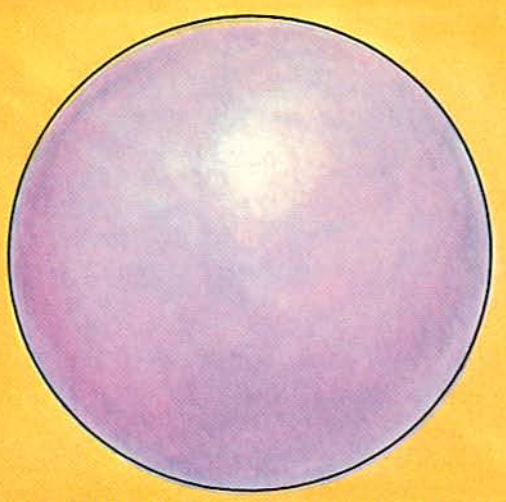
By then the French chemist, Joseph Louis Proust, had discovered that whenever elements form compounds these are always of a very definite composition — the "Law of Definite Composition." Water molecules, for example, always contain the same number of hydrogen and oxygen atoms. And Dalton had found that when two elements combine in different ways they do this in simple proportions — the "Law of Multiple Proportions." One atom of carbon and one atom of oxygen make carbon monoxide; one atom of carbon and two atoms of oxygen make carbon dioxide.

To describe these things in a simple way Berzelius made each of his symbols stand not only for a specific element but also for its relative weight as compared to the weight of other elements—its "atomic weight." To show the composition of a compound he simply put together the symbols for the elements into a "formula"—CO, HCl, FeS, and so on."CO" then not only meant that one atom of carbon and one atom of oxygen combine to make one molecule of carbon monoxide, but also that 12 weight units of carbon (12 being the atomic weight of carbon) combine with 16 weight units of oxygen (16 being the atomic weight of oxygen) to form 28 weight units of the compound carbon monoxide.

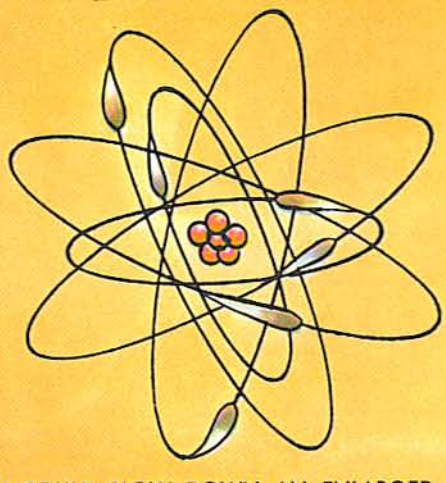
When a compound contained several atoms of the same element Berzelius indicated this by placing a number in front of the symbol. It was later found necessary to change this to a smaller number, called a "subscript," placed at the lower right of the symbol — H_2O, CO_2.

In recent years it has been necessary to change Dalton's idea of an atom as being the smallest indivisible part of an element. Nowadays we have machines, such as the cyclotron, that can bombard, or "smash" atoms into still smaller parts—neutrons, and electrically charged protons and electrons. According to today's atomic theory protons and neutrons form the nucleus of the atom and electrons whirl around the nucleus with such tremendous speed that they seem to form a "shell" around it.

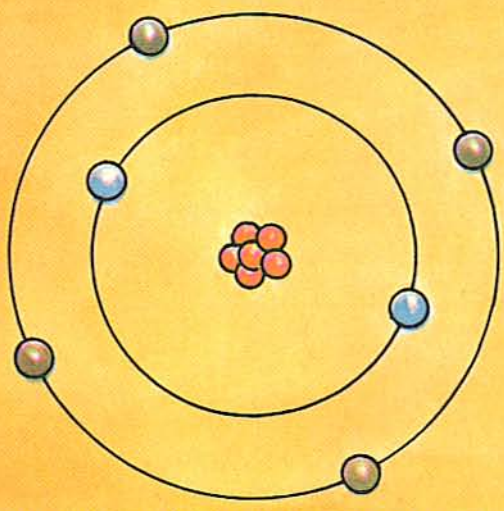
But even with our new idea of an atom, Dalton's main theory is still useful for explaining chemical reactions, and Berzelius' method is still the simplest "shorthand" method any scientist has ever devised for writing them down.



AN ATOM MIGHT LOOK LIKE A BALL SUCH AS THIS IF YOU ENLARGED IT A BILLION TIMES. THE "SHELL" IS NOT SOLID—IT CONSISTS OF ELECTRONS MOVING SO FAST THAT THEY SEEM TO FORM A SOLID SHELL.

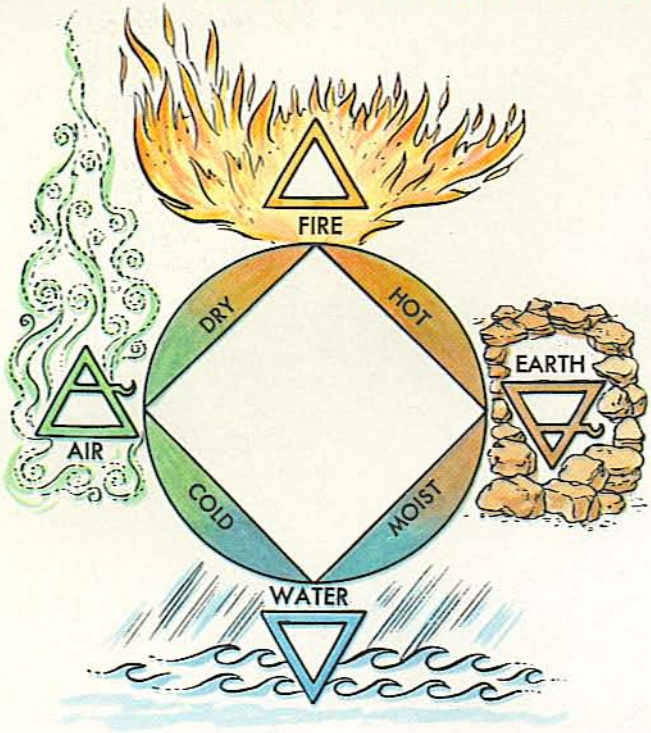


IF YOU COULD SLOW DOWN AN ENLARGED CARBON ATOM YOU MIGHT SEE TWO OF ITS ELECTRONS TRAVELING AROUND THE NUCLEUS IN AN "INNER SHELL" AND FOUR MORE WHIRLING AROUND IN AN "OUTER SHELL."



IF YOU COULD HALT AN ENLARGED CARBON ATOM COMPLETELY, IT WOULD LOOK A LOT LIKE OUR SOLAR SYSTEM, WITH A "SUN" (PROTONS AND NEUTRONS) IN THE CENTER AND "PLANETS" (ELECTRONS) AROUND IT.

The Periodic Table of the Elements

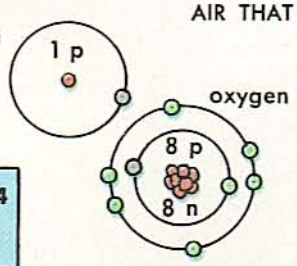


FROM THE earliest times people have tried to explain what "matter" was made of. Most early philosophers agreed that "matter" was made up of what they called "elements." But their idea of an "element" was quite different from what we mean by that word today.

The early Greek philosophers thought the entire universe was composed of only four basic substances: fire, earth, water, and air. This explanation made sense at the time and was not seriously challenged for many centuries.

The old Romans actually knew nine of the substances we call elements today. They called them, of course, by their Latin names (the same we use today in chemical symbols): *carbo* (carbon — C), *sulfur* (S), *aurum* (gold — Au), *argentum* (silver —

FOR MORE THAN A THOUSAND YEARS PHILOSOPHERS INSISTED THAT ALL SUBSTANCES WERE MADE UP OF FOUR ELEMENTS: FIRE THAT WAS DRY AND HOT, EARTH THAT WAS HOT AND MOIST, WATER THAT WAS MOIST AND COLD, AIR THAT WAS COLD AND DRY. WE KNOW BETTER NOW!



THE MODERN PICTURE OF AN ATOM HAS A NUCLEUS IN THE CENTER, CONSISTING OF PROTONS (p) AND NEUTRONS (n), WITH ELECTRONS IN RINGS AROUND IT.

THE PERIODIC TABLE OF THE ELEMENTS

	O	I A	II A	III A	IV A	V A	VI A	VII A	
1	H 1 Hydrogen 1.008								
2	He 2 Helium 4.003	Li 3 Lithium 6.940	Be 4 Beryllium 9.013						
3	Ne 10 Neon 20.183	Na 11 Sodium 22.991	Mg 12 Magnesium 24.32						
4	A 18 Argon 39.944	K 19 Potassium 39.1	Ca 20 Calcium 40.08	Sc 21 Scandium 44.96	Ti 22 Titanium 47.9	V 23 Vanadium 50.95	Cr 24 Chromium 52.01	Mn 25 Manganese 54.94	Fe 26 Iron 55.85
5	Kr 36 Krypton 83.8	Rb 37 Rubidium 85.48	Sr 38 Strontium 87.63	Y 39 Yttrium 88.92	Zr 40 Zirconium 91.22	Nb 41 Niobium 92.91	Mo 42 Molybdenum 95.95	Tc 43 Technetium 99	Ru 44 Ruthenium 101.1
6	Xe 54 Xenon 131.3	Cs 55 Cesium 132.91	Ba 56 Barium 137.36	57-71 Lanthanons	Hf 72 Hafnium 178.50	Ta 73 Tantalum 180.95	W 74 Tungsten 183.86	Re 75 Rhenium 186.22	Os 76 Osmium 190.2
7	Rn 86 Radon 222	Fr 87 Francium 223	Ra 88 Radium 226.05	89-103 Actinons					

La 57 Lanthanum 138.92	Ce 58 Cerium 140.13	Pr 59 Praseodymium 140.92	Nd 60 Neodymium 144.27	Pm 61 Promethium 145	Sm 62 Samarium 150.35
Ac 89 Actinium 227	Th 90 Thorium 232.05	Pa 91 Protactinium 231	U 92 Uranium 238.07	Np 93 Neptunium 237	Pu 94 Plutonium 242

INERT GASES — ALKALI METALS — ALKALINE EARTH METALS

ROWS RUNNING FROM LEFT TO RIGHT ARE CALLED PERIODS. COLUMNS RUNNING FROM TOP TO BOTTOM ARE CALLED GROUPS. THE ELEMENTS WITHIN A GROUP HAVE MANY TRAITS IN COMMON.

Ag), *ferrum* (iron — Fe), *cuprum* (copper — Cu), *stannum* (tin — Sn), *plumbum* (lead — Pb), *hydrargyrum* (mercury—Hg).

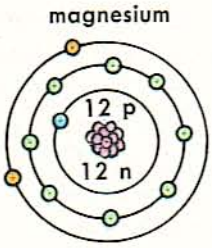
By 1800, thirty-four elements had been discovered. Within the next ten years, thirteen more had been added and had been given made-up Latin names — among them *natrium* (sodium — Na), *kalium* (potassium — K), and *aluminium* (aluminum — Al). By the beginning of the twentieth century, eighty-four elements were known.

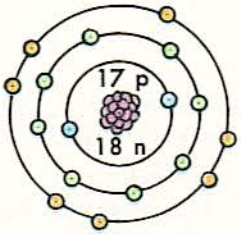
Today the number has reached 102 — the last ten man-made, produced by splitting the atoms of other elements. Within a short time, Element 103 will probably be discovered.

In this table you will find listed the 102 elements that are known today. Each element is described by its chemical symbol, its atomic number, its full name, and its atomic weight.

MANY SCIENTISTS HAD NOTICED THAT IF YOU LINE UP THE ELEMENTS ACCORDING TO ATOMIC WEIGHTS, CERTAIN CHEMICAL TRAITS OCCUR PERIODICALLY. THE RUSSIAN SCIENTIST, DMITRI MENDELEEFF, ON THIS BASIS DISCOVERED THE PERIODIC LAW AND DEVELOPED THE PERIODIC TABLE.

A YOUNG ENGLISH SCIENTIST, HENRY MOSELEY, PERFECTED THE PERIODIC TABLE. HE DISCOVERED THE LAW OF ATOMIC NUMBERS AND ARRANGED THE ELEMENTS ACCORDING TO THE ELECTRIC CHARGE FOUND IN THE NUCLEUS.





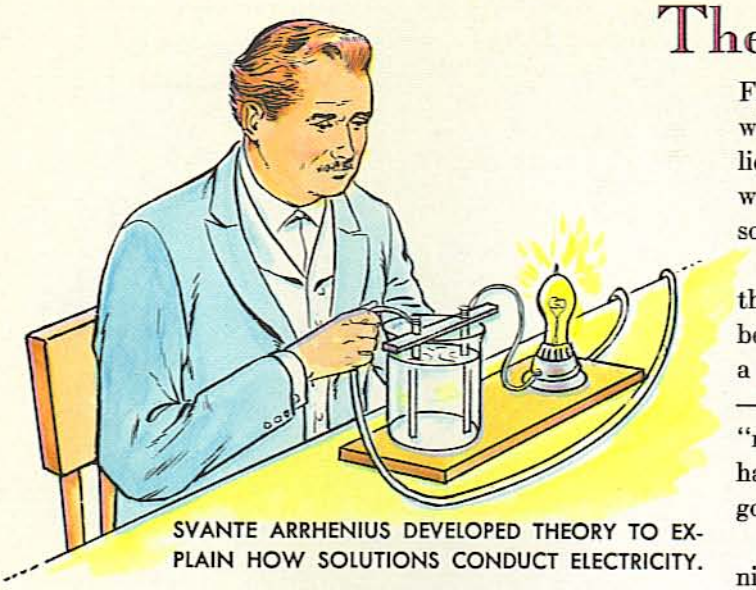
THE NUMBER OF PROTONS IN AN ATOM IS ITS ATOMIC NUMBER. AN ATOM ALWAYS HAS THE SAME NUMBER OF PROTONS AND ELECTRONS. HYDROGEN IS THE SIMPLEST OF ALL ATOMS.

VIII			I B	II B	III B	IV B	V B	VI B	VII B
					B 5 Boron 10.82	C 6 Carbon 12.011	N 7 Nitrogen 14.008	O 8 Oxygen 16	F 9 Fluorine 19
					Al 13 Aluminum 26.98	Si 14 Silicon 28.09	P 15 Phosphorus 30.975	S 16 Sulfur 32.066	Cl 17 Chlorine 35.457
Co 27 Cobalt 58.94	Ni 28 Nickel 58.71		Cu 29 Copper 63.54	Zn 30 Zinc 65.38	Ga 31 Gallium 69.72	Ge 32 Germanium 72.6	As 33 Arsenic 74.91	Se 34 Selenium 78.96	Br 35 Bromine 79.916
Rh 45 Rhodium 102.91	Pd 46 Palladium 106.4		Ag 47 Silver 107.88	Cd 48 Cadmium 112.41	In 49 Indium 114.82	Sn 50 Tin 118.7	Sb 51 Antimony 121.76	Te 52 Tellurium 127.61	I 53 Iodine 126.91
Ir 77 Iridium 192.2	Pt 78 Platinum 195.09		Au 79 Gold 197	Hg 80 Mercury 200.61	Tl 81 Thallium 204.39	Pb 82 Lead 207.21	Bi 83 Bismuth 209	Po 84 Polonium 210	At 85 Astatine 210

HEAVY METALS — METALLOIDS — NON-METALS — RARE EARTH METALS — UNSTABLE ELEMENTS

Eu 63 Europium 152	Gd 64 Gadolinium 157.26	Tb 65 Terbium 158.93	Dy 66 Dysprosium 162.51	Ho 67 Holmium 164.94	Er 68 Erbium 167.27	Tm 69 Thulium 168.94	Yb 70 Ytterbium 173.04	Lu 71 Lutetium 174.99
Am 95 Americium 243	Cm 96 Curium 247	Bk 97 Berkelium 249	Cf 98 Californium 249	E 99 Einsteinium 254	Fm 100 Fermium 255	Mv 101 Mendelevium 256	No 102 Nobelium 251	? 103

The Mysteries of Solutions



SVANTE ARRHENIUS DEVELOPED THEORY TO EXPLAIN HOW SOLUTIONS CONDUCT ELECTRICITY.

FROM THE earliest days, scientists experimenting with chemistry have worked with solutions. The liquid they used for making a solution (usually water) they called the "solvent." The chemical dissolved was the "solute."

When chemists began to use electricity as one of their tools, they discovered that different solutions behaved in different ways. The solution in water of a great number of chemicals — sugar among them — did not let electricity pass through. They were "non-conductors." Some chemicals, on the other hand, conducted electricity very easily. They were good conductors — "electrolytes."

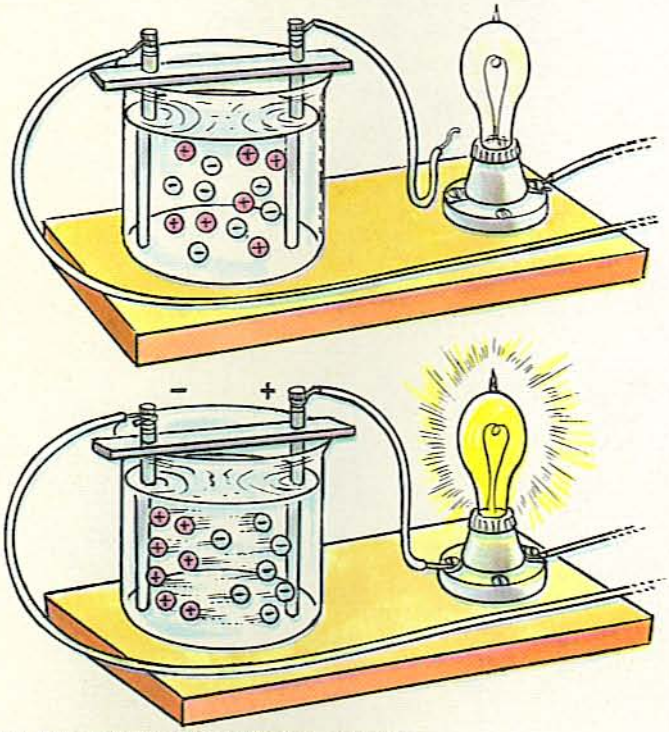
In 1874 a Swedish scientist named Svante Arrhenius developed a theory to help explain the mysterious behavior of solutions. He was only 25 years old at the time.

His idea was that when a chemical that conducts electricity is dissolved in water, each molecule is broken up — "dissociated" — into electrically charged atoms or groups of atoms. These atoms or groups of atoms Arrhenius called "ions" from a Greek word that means "to wander." His new theory came to be called "Arrhenius' theory of ionization."

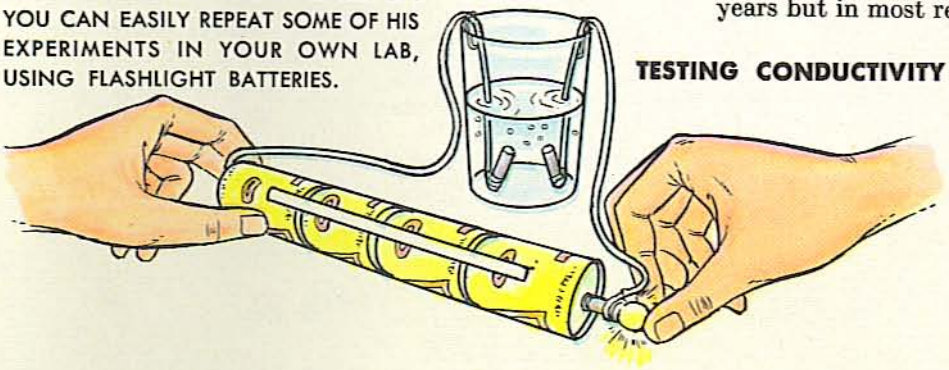
When table salt (sodium chloride, NaCl), for instance, is dissolved in water, it ionizes into positively charged sodium ions (Na^+) and negatively charged chlorine ions (Cl^-). These ions "wander" about in all directions until an electric current is applied to the solution. When that happens, the negative ions rush to the positive pole, the positive ions to the negative pole. It is the ions that conduct the current through the solution.

The reason that non-conductors do not conduct electricity is that they do not dissociate into ions.

Arrhenius' theory of ionization helped explain a great number of things that have puzzled chemists. His theory has been modified somewhat over the years but in most respects holds true today.



IN HIS EARLY EXPERIMENTS, SVANTE ARRHENIUS USED A SIMPLE SET-UP. YOU CAN EASILY REPEAT SOME OF HIS EXPERIMENTS IN YOUR OWN LAB, USING FLASHLIGHT BATTERIES.

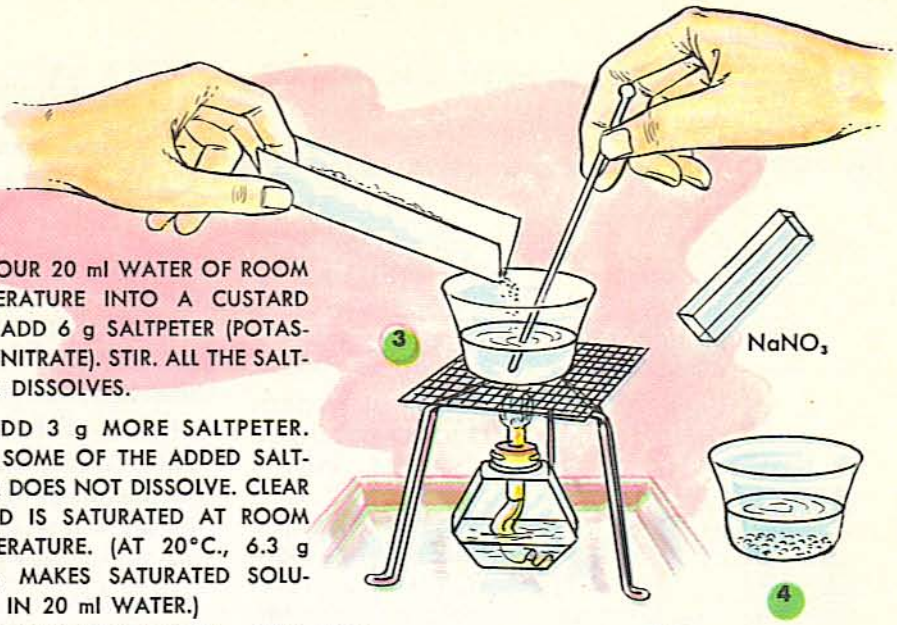


TESTING CONDUCTIVITY OF SOLUTIONS

SET UP THE SAME APPARATUS AS ON PAGE 25. ADD FLASHLIGHT BULB TO END OF ONE WIRE. TRY DIFFERENT SOLUTIONS IN GLASS. SOME CONDUCT ELECTRICITY AND BULB LIGHTS UP. OTHERS DO NOT AND THE BULB DOES NOT LIGHT UP.

SATURATED SOLUTIONS

A SATURATED SOLUTION IS ONE IN WHICH NO MORE OF THE CHEMICAL WILL GO IN SOLUTION AT THAT PARTICULAR TEMPERATURE.



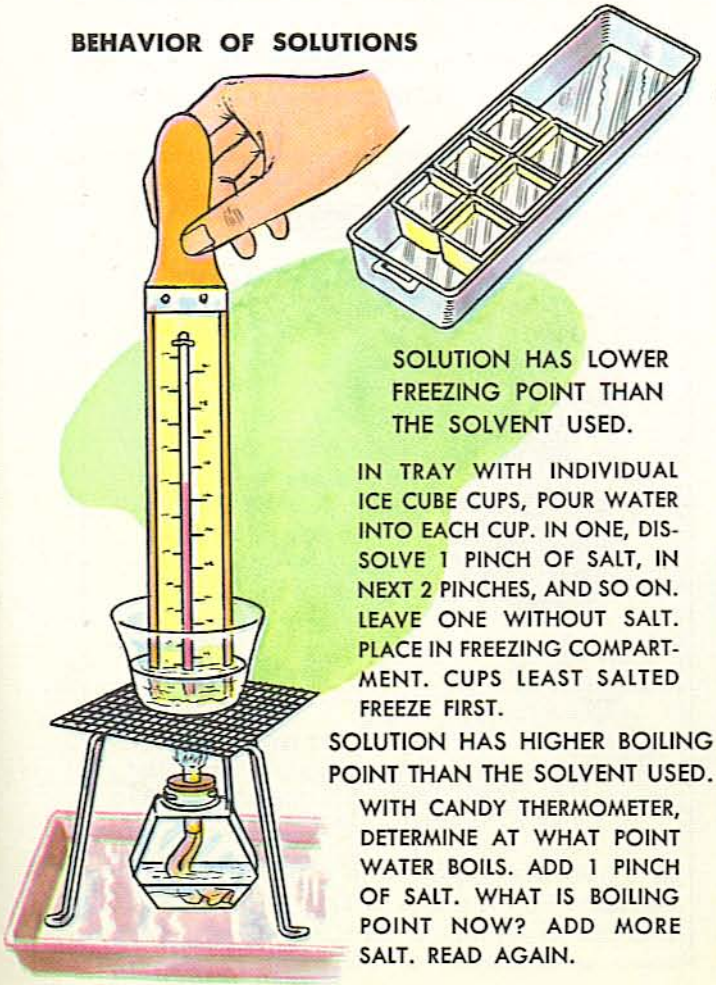
1 POUR 20 ml WATER OF ROOM TEMPERATURE INTO A CUSTARD CUP. ADD 6 g SALTPETER (POTASSIUM NITRATE). STIR. ALL THE SALTPETER DISSOLVES.

2 ADD 3 g MORE SALTPETER. STIR. SOME OF THE ADDED SALTPETER DOES NOT DISSOLVE. CLEAR LIQUID IS SATURATED AT ROOM TEMPERATURE. (AT 20°C., 6.3 g KNO_3 MAKES SATURATED SOLUTION IN 20 ml WATER.)

3 PLACE CUSTARD CUP OVER ALCOHOL BURNER. ADD 10 g MORE SALTPETER. SOON ALL SALTPETER IS DISSOLVED. AT HIGHER TEMPERATURES IT TAKES MORE SOLUTE TO MAKE A SATURATED SOLUTION. (AT BOILING, 20 ml H_2O DISSOLVES 49 g SALTPETER.)

4 TAKE SOLUTION OFF FIRE. AS IT COOLS, MUCH OF THE SALTPETER COMES OUT AS CRYSTALS BY SLOW CRYSTALLIZATION. LIQUID IS AGAIN A SOLUTION SATURATED AT ROOM TEMPERATURE.

BEHAVIOR OF SOLUTIONS



SOLUTION HAS LOWER FREEZING POINT THAN THE SOLVENT USED.

IN TRAY WITH INDIVIDUAL ICE CUBE CUPS, POUR WATER INTO EACH CUP. IN ONE, DISSOLVE 1 PINCH OF SALT, IN NEXT 2 PINCHES, AND SO ON. LEAVE ONE WITHOUT SALT. PLACE IN FREEZING COMPARTMENT. CUPS LEAST SALTED FREEZE FIRST.

SOLUTION HAS HIGHER BOILING POINT THAN THE SOLVENT USED.

WITH CANDY THERMOMETER, DETERMINE AT WHAT POINT WATER BOILS. ADD 1 PINCH OF SALT. WHAT IS BOILING POINT NOW? ADD MORE SALT. READ AGAIN.

CRYSTALLIZATION

YOU CAN FOLLOW CRYSTALLIZATION OF $MgSO_4$. IN TEST TUBE, HEAT MIXTURE OF 5 ml WATER AND 1 TEASPOON EPSOM SALT UNTIL SALT DISSOLVES. POUR HOT SOLUTION OVER PANE OF GLASS CLEANED WITH DETERGENT. CRYSTALS MAKE NEEDLE-LIKE NETWORK.

MANY CHEMICALS FORM CRYSTALS OF DISTINCT SHAPES.

MAKING SOLUTIONS

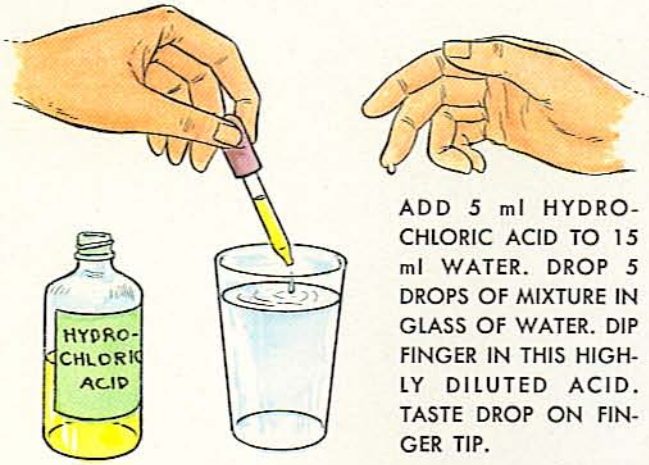
MAKE 50 ml GRADUATE FIRST: MEASURE 50 ml WATER INTO A NARROW JAR, USING 10 ml TEST TUBE GRADUATE SHOWN ON PAGE 15. MAKE A MARK AT 50 ml LEVEL.

10% (10 PER CENT) SOLUTION: MEASURE 40 ml WATER INTO A CUSTARD CUP. ADD 5 g OF THE CHEMICAL. STIR. (TO MAKE IT DISSOLVE QUICKER, YOU MAY WANT TO HEAT THE WATER SLIGHTLY.) POUR SOLUTION INTO 50 ml GRADUATE. ADD WATER TO THE 50 ml MARK.

2% SOLUTION: MEASURE 40 ml WATER INTO CUSTARD CUP. ADD 1 g OF THE CHEMICAL. STIR TO DISSOLVE. POUR INTO 50 ml GRADUATE. ADD WATER TO 50 ml.

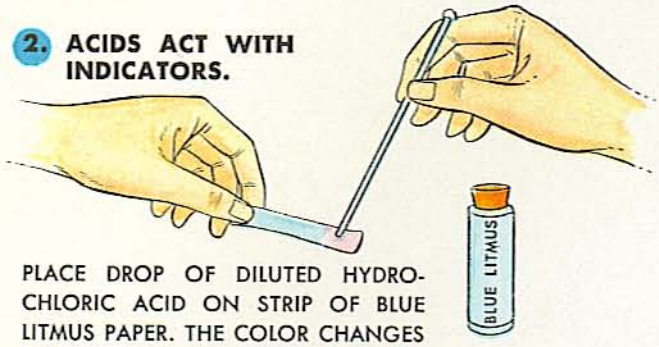
HOW DO YOU KNOW AN ACID?

1. ACIDS TASTE SOUR.



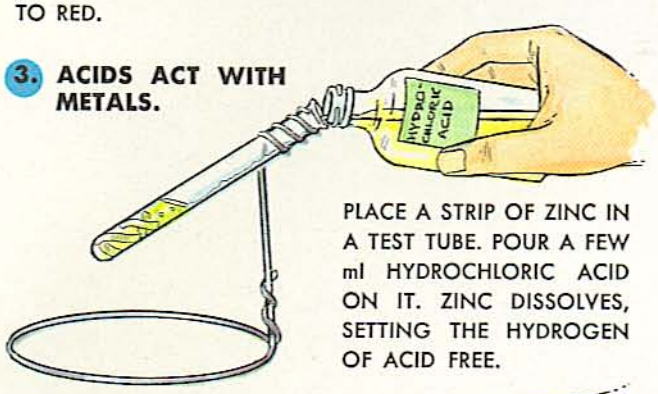
ADD 5 ml HYDROCHLORIC ACID TO 15 ml WATER. DROP 5 DROPS OF MIXTURE IN GLASS OF WATER. DIP FINGER IN THIS HIGHLY DILUTED ACID. TASTE DROP ON FINGER TIP.

2. ACIDS ACT WITH INDICATORS.



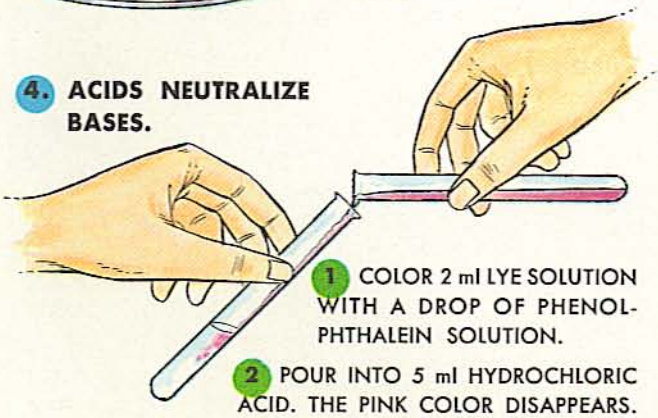
PLACE DROP OF DILUTED HYDROCHLORIC ACID ON STRIP OF BLUE LITMUS PAPER. THE COLOR CHANGES TO RED.

3. ACIDS ACT WITH METALS.



PLACE A STRIP OF ZINC IN A TEST TUBE. POUR A FEW ml HYDROCHLORIC ACID ON IT. ZINC DISSOLVES, SETTING THE HYDROGEN OF ACID FREE.

4. ACIDS NEUTRALIZE BASES.



1 COLOR 2 ml LYE SOLUTION WITH A DROP OF PHENOLPHTHALEIN SOLUTION.

2 POUR INTO 5 ml HYDROCHLORIC ACID. THE PINK COLOR DISAPPEARS.

Working With Acids

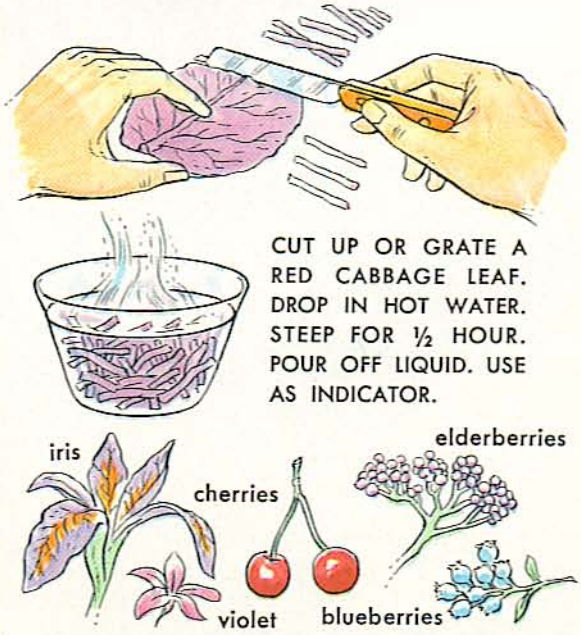
ACIDS have many traits in common. They taste sour. They change the color of certain plant substances—which are called "indicators." They contain hydrogen (H) that can be replaced by a metal. They neutralize bases.

But what *is* an acid? Earlier, the "acidic" traits were used to define an acid. But with the modern understanding of the atom, a different definition is used. You will remember that the nucleus of an atom contains positively charged protons. Acids in solution liberate protons as ions (H^+). And so we say that an acid is a substance that will give up — or "donate" — protons to another substance. Acids are "proton donors." The foremost acids used in industry are sulfuric acid (H_2SO_4), nitric acid (HNO_3), and hydrochloric acid (HCl).

The first two — sulfuric acid and nitric acid — should NEVER be used in the home lab. They are much too DANGEROUS. They destroy the skin and might blind you if you got them in the eyes. (Wherever a chemical experiment would ordinarily call for sulfuric acid, this book uses sodium acid sulfate — $NaHSO_4$, sodium bisulfate, "Sani-Flush"; wherever

(CONTINUED ON PAGE 44)

HOME-MADE INDICATORS



CUT UP OR GRATE A RED CABBAGE LEAF. DROP IN HOT WATER. STEEP FOR ½ HOUR. POUR OFF LIQUID. USE AS INDICATOR.

MANY FLOWERS AND FRUITS CONTAIN COLORING MATTER WHICH YOU CAN EXTRACT WITH HOT WATER AND USE AS AN INDICATOR FOR ACIDS AND BASES.

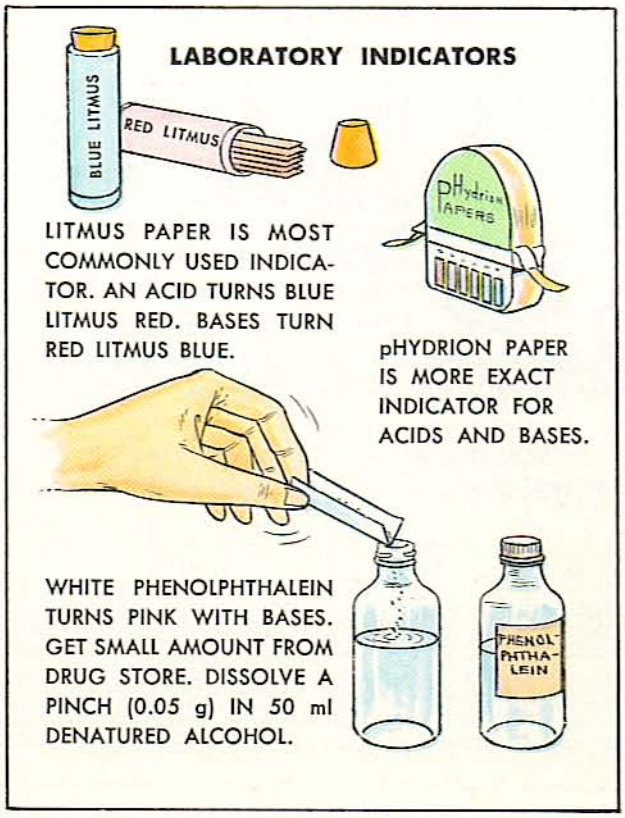
Working With Bases

BASES taste brackish. They change the color of "indicators." They contain a combination of oxygen and hydrogen atoms called "hydroxyl" (OH). They neutralize acids.

But what *is* a base? When a base is dissolved in water it liberates negatively charged hydroxyl ions (OH^-). When a base is neutralized, these ions take on — or "accept"— positively charged protons from another substance. A base is a substance that will accept and combine with protons from another substance. Bases are "proton acceptors." The most important bases are sodium hydroxide ("lye," NaOH), ammonium hydroxide ("ammonia," NH_4OH), and calcium hydroxide ("slaked lime," $Ca(OH)_2$).

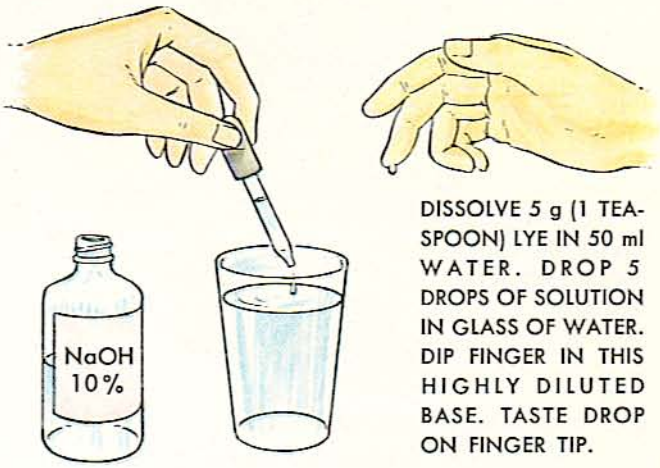
The first of these — sodium hydroxide — is used in many households to clean sluggish drains and to keep sinks from stopping up ("Drano"). USE IT WITH GREAT CARE in your experiments. Do not touch lye flakes with your fingers and do not get the solution on your skin — it dissolves the natural oil. It is particularly dangerous to get lye in your eyes. If you get lye on you, dilute it quickly with LOTS OF WATER.

(CONTINUED ON PAGE 45)



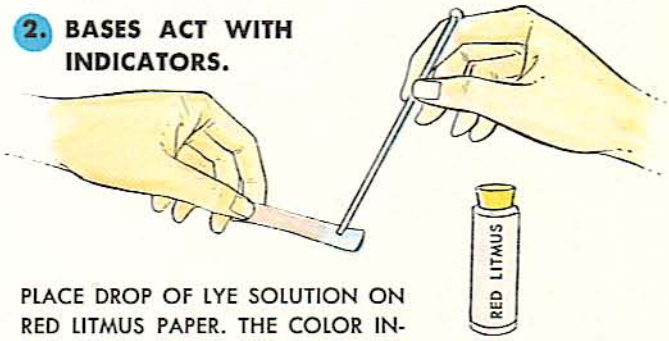
HOW DO YOU KNOW A BASE?

1. BASES TASTE BRACKISH.



DISSOLVE 5 g (1 TEASPOON) LYE IN 50 ml WATER. DROP 5 DROPS OF SOLUTION IN GLASS OF WATER. DIP FINGER IN THIS HIGHLY DILUTED BASE. TASTE DROP ON FINGER TIP.

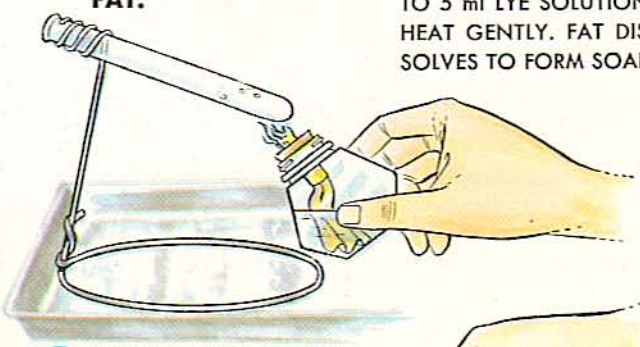
2. BASES ACT WITH INDICATORS.



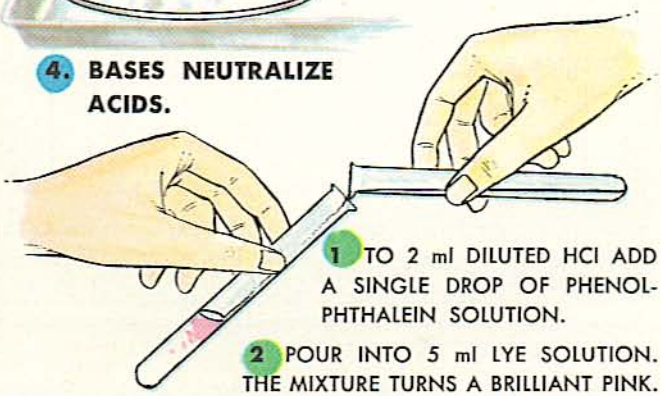
PLACE DROP OF LYE SOLUTION ON RED LITMUS PAPER. THE COLOR INSTANTLY CHANGES TO BLUE.

3. BASES ACT WITH FAT.

ADD TINY LUMP OF FAT TO 5 ml LYE SOLUTION. HEAT GENTLY. FAT DISSOLVES TO FORM SOAP.



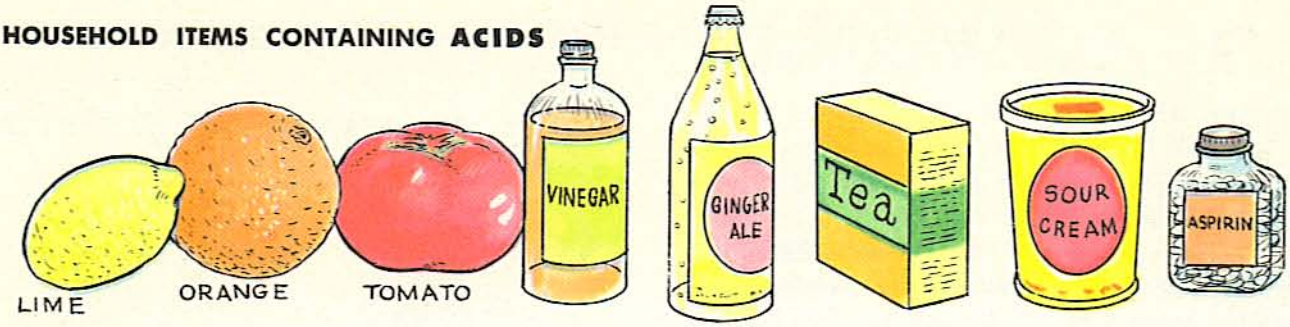
4. BASES NEUTRALIZE ACIDS.



1 TO 2 ml DILUTED HCl ADD A SINGLE DROP OF PHENOLPHTHALEIN SOLUTION.

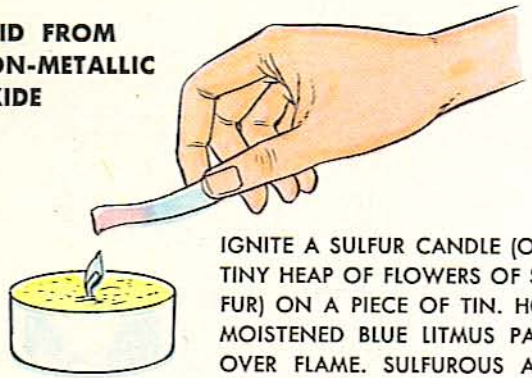
2 POUR INTO 5 ml LYE SOLUTION. THE MIXTURE TURNS A BRILLIANT PINK.

HOUSEHOLD ITEMS CONTAINING ACIDS



Acids—Continued

ACID FROM NON-METALLIC OXIDE

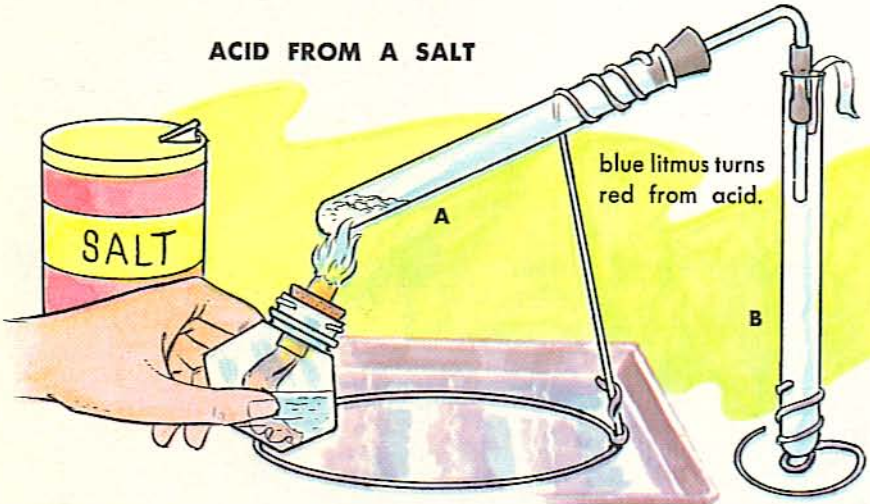


IGNITE A SULFUR CANDLE (OR A TINY HEAP OF FLOWERS OF SULFUR) ON A PIECE OF TIN. HOLD MOISTENED BLUE LITMUS PAPER OVER FLAME. SULFUROUS ACID FORMED TURNS IT RED.

nitric acid would be called for, this book produces it in a mixture of a nitrate, KNO_3, and sodium bisulfate.)

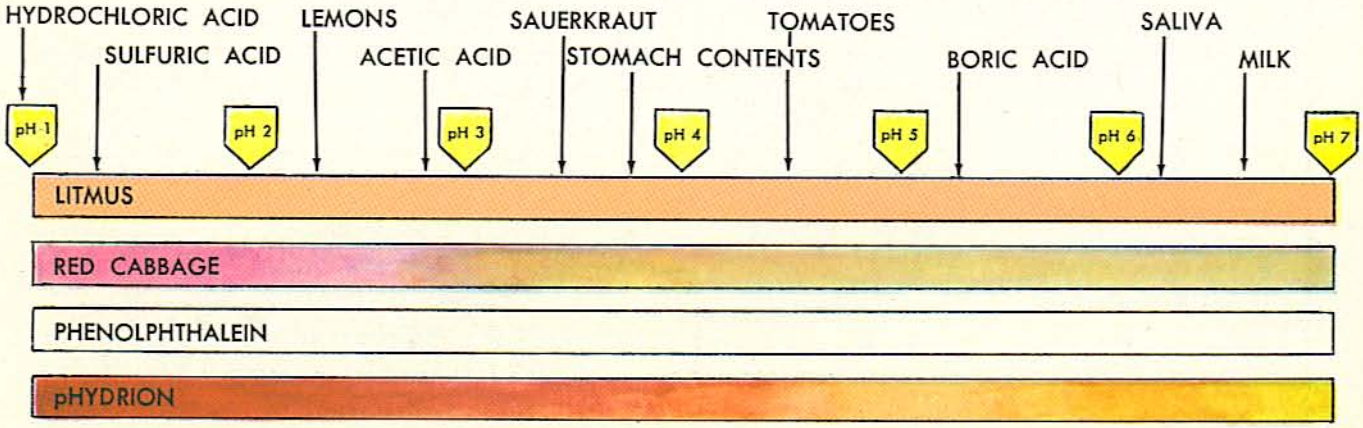
Hydrochloric acid is used in many households under the name of "muriatic acid." Whenever you use hydrochloric acid in an experiment; USE IT WITH GREAT CARE. If any of it gets on you, dilute it quickly with LOTS OF WATER. Or neutralize it with bicarbonate of soda (but not if in the eyes).

ACID FROM A SALT

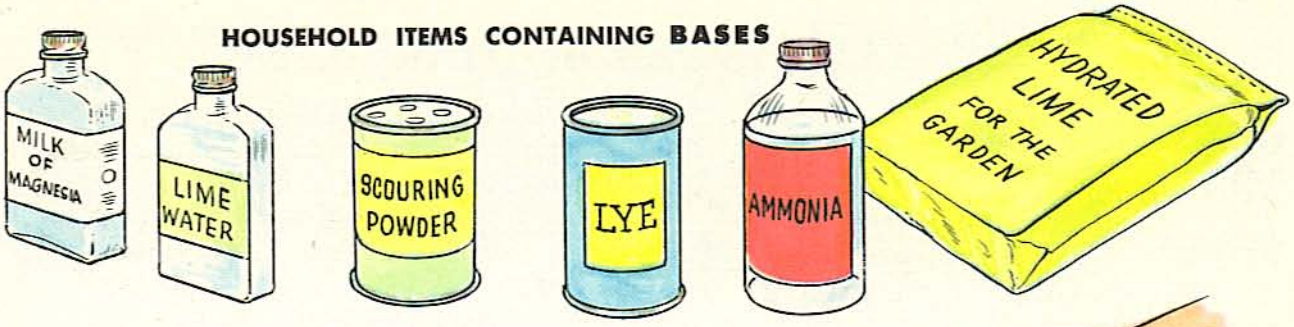


SET UP APPARATUS AS SHOWN. INTO TEST TUBE **A** DROP MIXTURE OF ¼ TEASPOON TABLE SALT AND ½ TEASPOON SODIUM BISULFATE. HEAT. HYDROGEN CHLORIDE PRODUCED TURNS MOISTENED BLUE LITMUS RED. ADD 2 ml WATER TO TEST TUBE **B**. SHAKE. RESULT IS WEAK HYDROCHLORIC ACID.

pH SYSTEM IS A WAY OF DESCRIBING THE RELATIVE ACIDITY OR ALKALINITY OF A SOLUTION. PURE WATER IS NEUTRAL WITH pH7. THE LOWER THE NUMBER BELOW 7, THE MORE ACID THE SOLUTION. THE HIGHER



HOUSEHOLD ITEMS CONTAINING BASES



Bases—Continued

You can also neutralize it with vinegar (but not if in the eyes).

Ammonia is a common household cleaning liquid. Ammonia should also be handled with care and should be washed off quickly if you get it on you. Also watch your nose when you work with ammonia. It has a very strong smell.

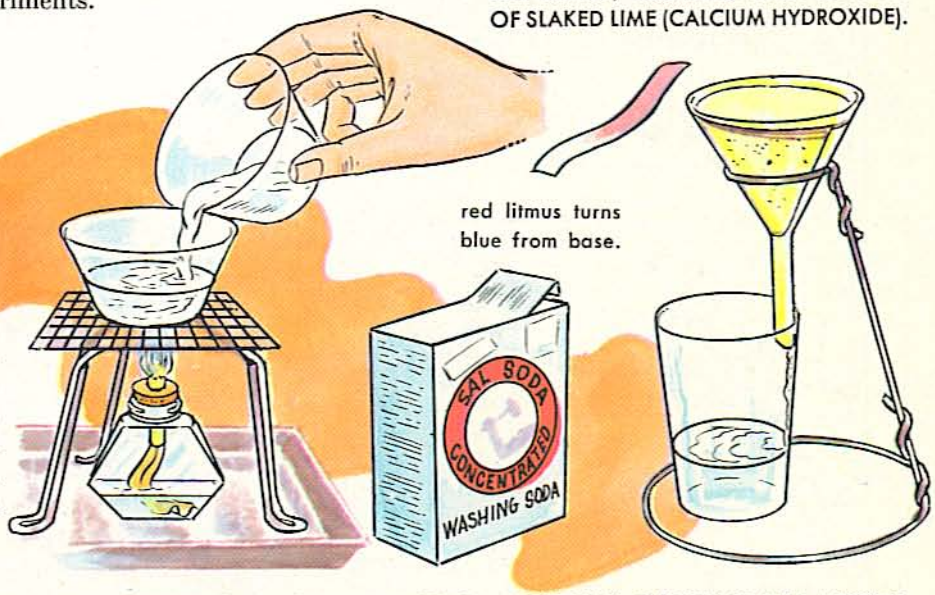
Calcium hydroxide is a white powder. You will use it in a great number of experiments.

BASE FROM METALLIC OXIDE

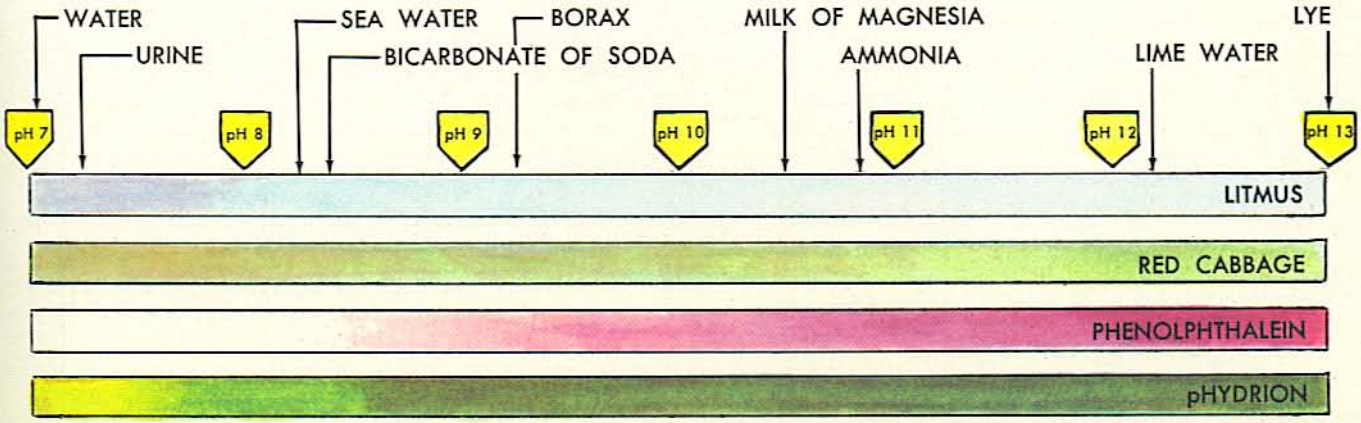
PLACE A LUMP OF LIME (QUICKLIME, CALCIUM OXIDE) IN A CUSTARD CUP. ADD AS MUCH LUKEWARM WATER AS IT WILL ABSORB. LIME HEATS UP, GIVES OFF STEAM, CRUMBLES INTO POWDER OF SLAKED LIME (CALCIUM HYDROXIDE).

BASE FROM A SALT

IN A CUSTARD CUP, DISSOLVE 1 TEASPOON SAL SODA (WASHING SODA, SODIUM CARBONATE) IN 50 ml WATER. HEAT SLIGHTLY. ADD SLAKED LIME MIXED WITH WATER. STIR. CHEMICAL REACTION PRODUCES SODIUM HYDROXIDE AND CALCIUM CARBONATE. FILTER. CLEAR LIQUID CONTAINS THE SODIUM HYDROXIDE (LYE). THE CALCIUM CARBONATE IS HELD BACK BY THE FILTER.



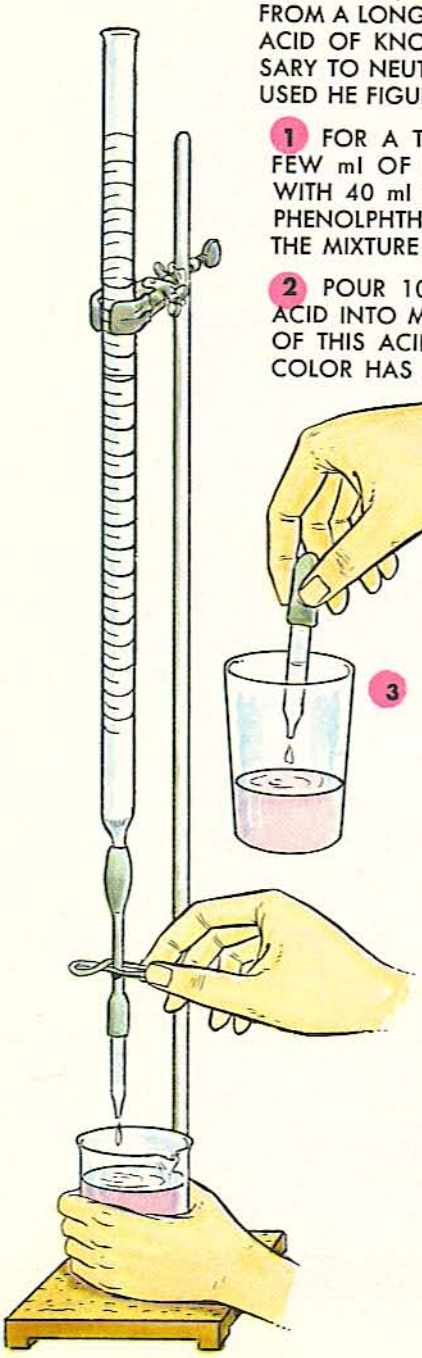
THE NUMBER ABOVE 7, THE MORE ALKALINE THE SOLUTION. WHEN YOU KNOW AT WHAT pH AN INDICATOR CHANGES COLOR, YOU CAN DETERMINE THE ACIDITY OR ALKALINITY OF THE SOLUTION YOU ARE TESTING.



Salts—Chemicals of Many Uses

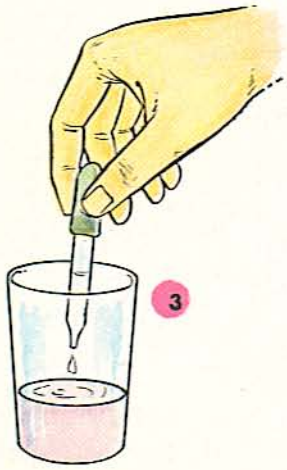
NEUTRALIZATION IS USED EXTENSIVELY IN CHEMICAL ANALYSIS IN A TECHNIQUE CALLED **TITRATION.**

TO DETERMINE THE UNKNOWN STRENGTH OF A BASE, THE CHEMIST DROPS INTO IT FROM A LONG TUBE—A BURETTE—AS MUCH ACID OF KNOWN STRENGTH AS IS NECESSARY TO NEUTRALIZE IT. BY CHECKING ACID USED HE FIGURES STRENGTH OF BASE.



1 FOR A TRY AT TITRATION, MIX A FEW ml OF HOUSEHOLD AMMONIA WITH 40 ml WATER. ADD A DROP OF PHENOLPHTHALEIN. THIS WILL COLOR THE MIXTURE A DEEP PINK.

2 POUR 10 ml DILUTED HYDROCHLORIC ACID INTO MEASURING TUBE. POUR SOME OF THIS ACID INTO THE AMMONIA UNTIL COLOR HAS ALMOST VANISHED.



3 PICK UP A FEW ml OF THE MEASURED ACID IN AN EYE DROPPER (PIPETTE). DROP ACID SLOWLY INTO THE AMMONIA MIXTURE UNTIL COLOR IS COMPLETELY GONE. RETURN ACID NOT USED TO MEASURING TUBE. YOU NOW KNOW HOW MANY ml ACID YOU HAD TO USE TO NEUTRALIZE THE AMMONIA.

WHAT HAPPENS when you neutralize an acid with a base or a base with an acid? The hydrogen atoms (H^+ ions) of the acid combine with the hydroxyl groups (OH^- ions) of the base to form water, and the metal atoms of the base combine with what remains of the acid to form a salt. Or simply:

BASE plus ACID turns into
WATER plus SALT

This, for example, is what happens when you neutralize sodium hydroxide with hydrochloric acid:

$$NaOH + HCl \rightarrow HOH + NaCl$$

The result is water and sodium chloride — ordinary table salt which has given its name to other substances of a similar nature.

Of all the salts used in industry, table salt (NaCl) and washing soda (Na_2CO_3) are of greatest importance. Numerous other chemicals are produced from them. Our way of life would be completely disrupted if our country's industry did not have enough of these two salts.

Many other salts are necessary for our well-being. You'll probably find at least half a dozen different salts used daily in your home — in cooking and baking, in gardening, for cleaning.

In your chemical experiments you'll be working with two classes of salts: *normal salts* (such as NaCl, Na_2CO_3, KI) which contain no free hydrogen or hydroxyl ions, and *acid salts* (such as $NaHSO_4$, $NaHCO_3$) which contain replaceable hydrogen.

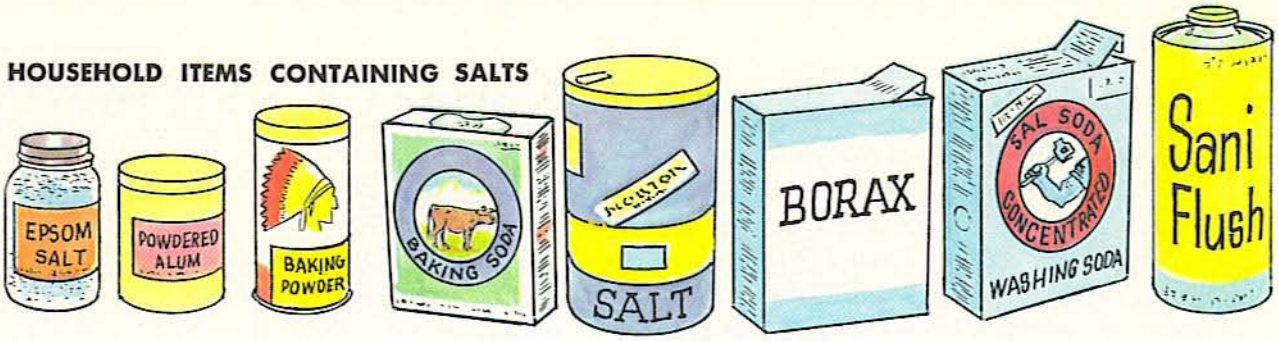
Some of these salts dissolve easily in water — all the nitrates (salts of nitric acid) and most of the chlorides (salts of hydrochloric acid). Many salts, on the other hand, are insoluble — most of the carbonates (salts of carbonic acid) and most sulfides (salts of hydrosulfuric acid).

HOW THE NAMES OF SALTS ARE MADE UP

THE ACID	FORMULA AND NAME OF ACID		FORMULA AND NAME OF SALT	
SULFURIC ACID	H_2SO_4	HYDROGEN SULFATE	Na_2SO_4	SODIUM SULFATE
NITRIC ACID	HNO_3	HYDROGEN NITRATE	$NaNO_3$	SODIUM NITRATE
CARBONIC ACID	H_2CO_3	HYDROGEN CARBONATE	Na_2CO_3	SODIUM CARBONATE
ACETIC ACID	$HC_2H_3O_2$	HYDROGEN ACETATE	$NaC_2H_3O_2$	SODIUM ACETATE
HYDROCHLORIC ACID	HCl	HYDROGEN CHLORIDE	NaCl	SODIUM CHLORIDE
HYDROSULFURIC ACID	H_2S	HYDROGEN SULFIDE	Na_2S	SODIUM SULFIDE
SULFUROUS ACID	H_2SO_3	HYDROGEN SULFITE	Na_2SO_3	SODIUM SULFITE
NITROUS ACID	HNO_2	HYDROGEN NITRITE	$NaNO_2$	SODIUM NITRITE
CHLOROUS ACID	$HClO_2$	HYDROGEN CHLORITE	$NaClO_2$	SODIUM CHLORITE

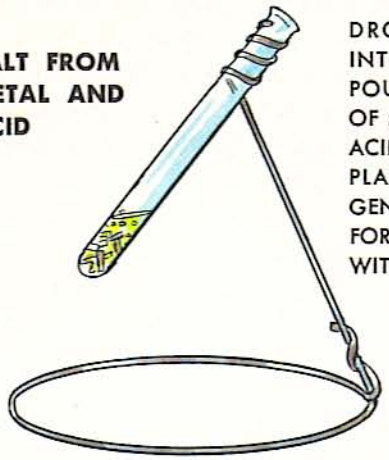
REMEMBER: **-IC** ACIDS FORM **-ATE** SALTS;
HYDRO- -IC ACIDS FORM **-IDE** SALTS; **-OUS** ACIDS FORM **-ITE** SALTS

HOUSEHOLD ITEMS CONTAINING SALTS



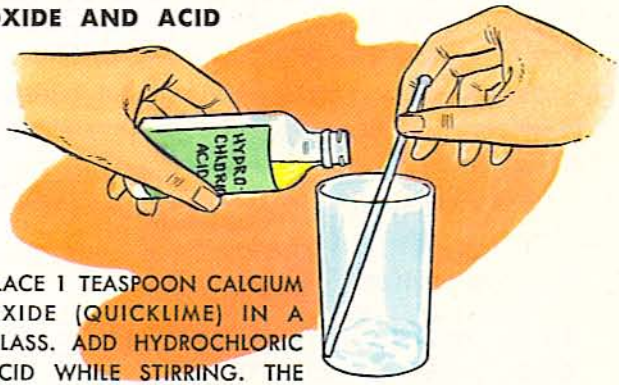
DIFFERENT WAYS OF PRODUCING SALTS

SALT FROM METAL AND ACID



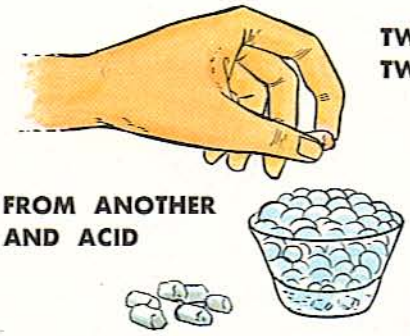
DROP ZINC STRIPS INTO A TEST TUBE. POUR IN A COUPLE OF ml HYDROCHLORIC ACID. THE ZINC DISPLACES THE HYDROGEN OF THE ACID TO FORM A SALT ($ZnCl_2$) WITH THE CHLORINE.

SALT FROM METAL OXIDE AND ACID



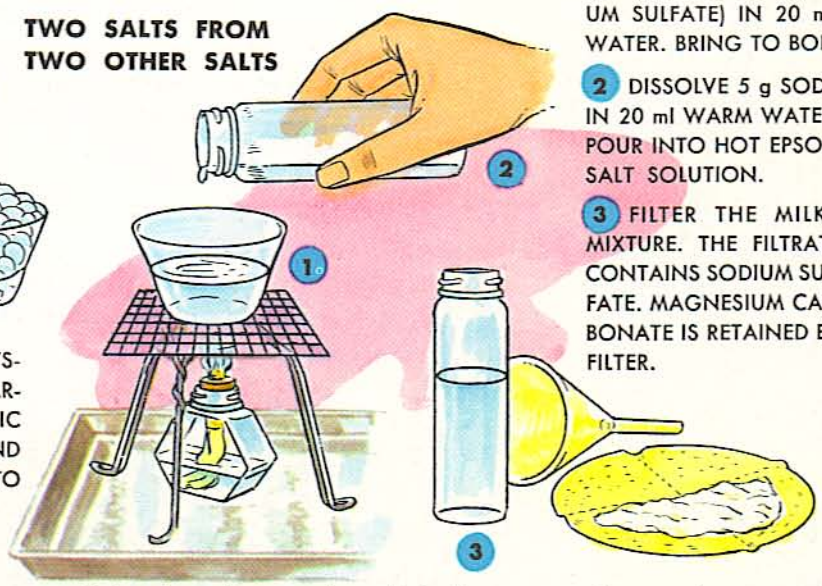
PLACE 1 TEASPOON CALCIUM OXIDE (QUICKLIME) IN A GLASS. ADD HYDROCHLORIC ACID WHILE STIRRING. THE QUICKLIME DISSOLVES IN THE ACID, FORMING CALCIUM CHLORIDE AND WATER.

SALT FROM ANOTHER SALT AND ACID



DROP PIECES OF CHALK, MARBLE, OR OYSTER SHELLS (ALL OF THEM CALCIUM CARBONATES) IN A FEW ml HYDROCHLORIC ACID. RESULT IS CALCIUM CHLORIDE AND CARBONIC ACID (WHICH BREAKS UP INTO CARBON DIOXIDE AND WATER).

TWO SALTS FROM TWO OTHER SALTS



1. DISSOLVE 5 g EPSOM SALTS (MAGNESIUM SULFATE) IN 20 ml. WATER. BRING TO BOIL.
2. DISSOLVE 5 g SODA IN 20 ml WARM WATER. POUR INTO HOT EPSOM SALT SOLUTION.
3. FILTER THE MILKY MIXTURE. THE FILTRATE CONTAINS SODIUM SULFATE. MAGNESIUM CARBONATE IS RETAINED BY FILTER.

THERE ARE many ways of producing a salt in addition to neutralization.

When you made iron sulfide directly from the two elements iron and sulfur, you produced a salt:

$$Fe + S \rightarrow FeS$$

When you caused zinc metal to react with hydrochloric acid, you made a salt:

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

When you made sodium hydroxide, you used a base and a salt to form a new base and a new salt:

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

A salt and an acid often form another salt and another acid:

$$CaCO_3 + 2HCl \rightarrow$$
$$CaCl_2 + H_2CO_3 \quad (H_2O + CO_2)$$

Two soluble salts may also form two other salts—one of them insoluble:

$$Na_2CO_3 + MgSO_4 \rightarrow MgCO_3 + Na_2SO_4$$

Iodine—Violet or Brown?

IODINE
Element 53.
At. wt. 126.91.
Gray-black crystals of a peculiar odor. Sublimes with violet color. Combines directly with metals and nonmetals. It has a density of 4.9.

IODINE IS an interesting element to experiment with. It is easily driven out of its compounds as beautiful, violet fumes that turn into grayish-black, metallic-looking crystals on cooling. These crystals can be further purified by turning them into vapor again, and again cooling them into crystal form. This process is called "sublimation."

You are probably familiar with the 2% alcoholic solution of iodine known as "tincture of iodine." It is found in almost every home medicine cabinet and is used as a disinfectant for wounds. Iodine has many other uses — in photography and in the preparation of various medicines and dyes.

Iodine has the bad habit of staining practically everything with which it comes in contact with a brown stain that won't come off in washing. That's why it is advisable to have sodium thiosulfate — photographer's fixing salt, "hypo"— around when you work with iodine. Hypo in solution forms a colorless compound with iodine.

Be careful not to breathe fumes.

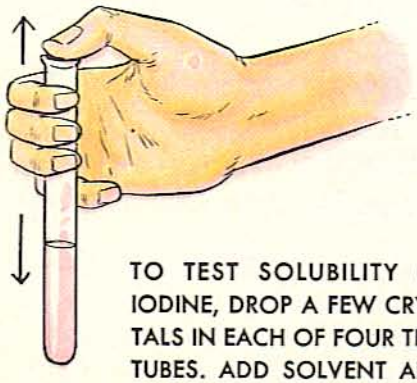
MAKING IODINE

1 IN A PYREX CUSTARD CUP MIX TOGETHER 2 g POTASSIUM IODIDE, 2 g MANGANESE DIOXIDE, 4 g SODIUM BISULFATE. HEAT MIXTURE GENTLY. SOON VIOLET FUMES EMERGE.

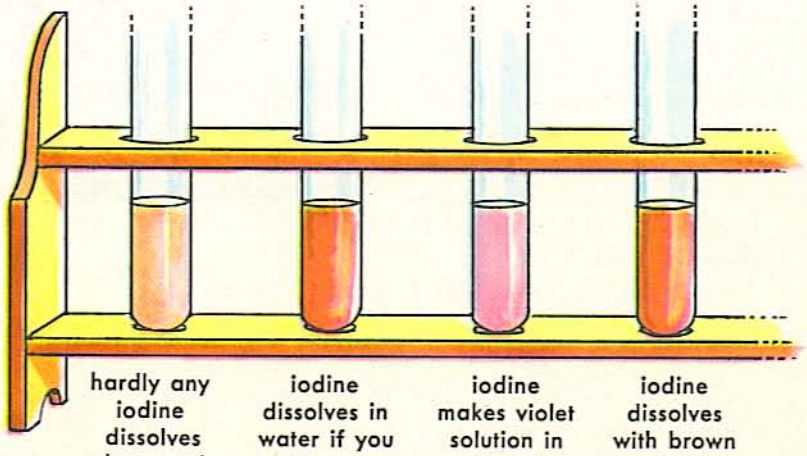
2 DROP HALF A DOZEN ICE CUBES INTO A JAR. ADD A LITTLE WATER. PLACE JAR AS A LID ON TOP OF CUSTARD CUP. THE VIOLET FUMES SETTLE ON BOTTOM OF JAR AS GRAYISH-BLACK, SHINY IODINE CRYSTALS.

3 SCRAPE IODINE CRYSTALS OFF BOTTOM OF JAR. KEEP THEM IN SMALL, TIGHTLY CLOSED BOTTLE.

SOLUBILITY OF IODINE



TO TEST SOLUBILITY OF IODINE, DROP A FEW CRYSTALS IN EACH OF FOUR TEST TUBES. ADD SOLVENT AND SHAKE TUBE.



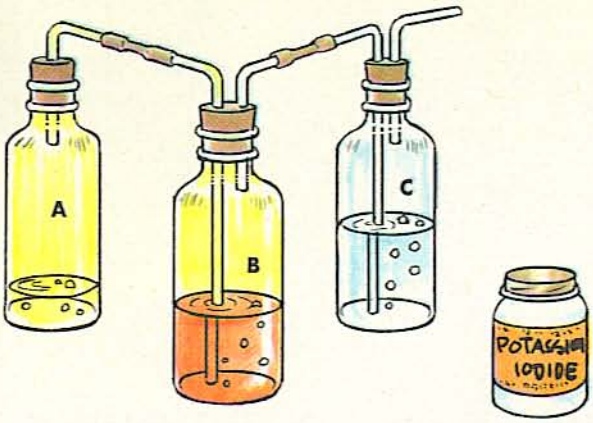
hardly any iodine dissolves when put in plain water.

iodine dissolves in water if you add potassium iodide.

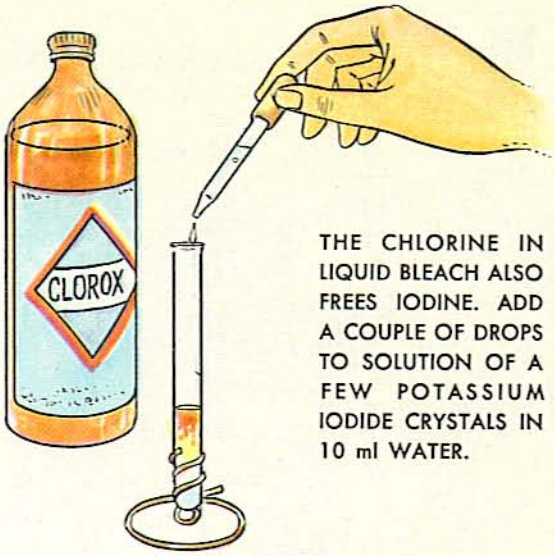
iodine makes violet solution in carbon tetrachloride.

iodine dissolves with brown color in alcohol.

IODINE FREED BY CHLORINE

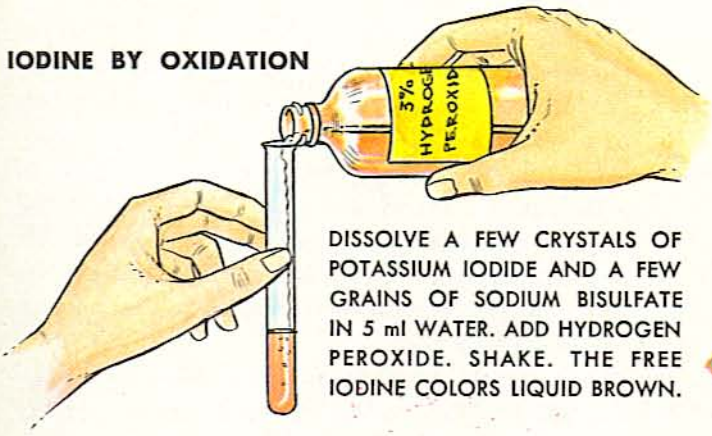


SET UP APPARATUS AS DESCRIBED ON PAGE 35 WITH THIS EXCEPTION: IN BOTTLE **B**, MAKE SOLUTION OF ½ g POTASSIUM IODIDE IN 40 ml WATER. AS CHLORINE BUBBLES THROUGH THIS SOLUTION IT TURNS BROWN FROM THE FREED IODINE. WITH MORE CHLORINE IT CLEARS AGAIN WHEN COLORLESS IODIC ACID FORMS.



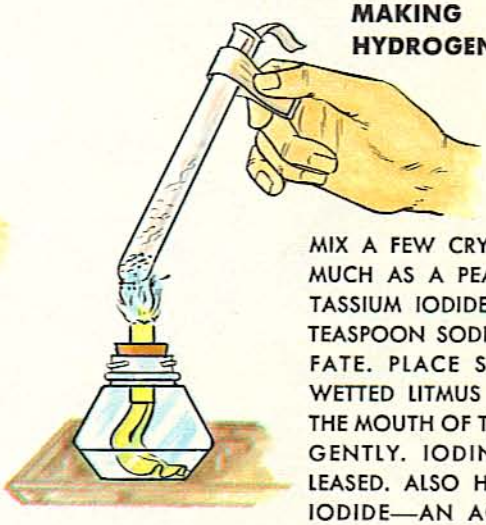
THE CHLORINE IN LIQUID BLEACH ALSO FREES IODINE. ADD A COUPLE OF DROPS TO SOLUTION OF A FEW POTASSIUM IODIDE CRYSTALS IN 10 ml WATER.

IODINE BY OXIDATION



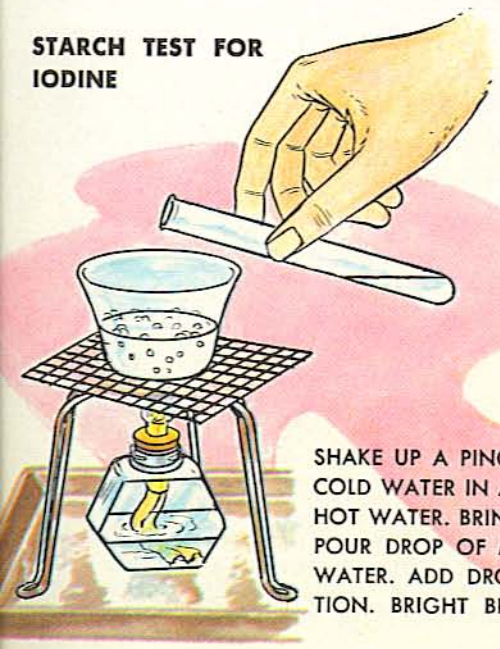
DISSOLVE A FEW CRYSTALS OF POTASSIUM IODIDE AND A FEW GRAINS OF SODIUM BISULFATE IN 5 ml WATER. ADD HYDROGEN PEROXIDE. SHAKE. THE FREE IODINE COLORS LIQUID BROWN.

MAKING HYDROGEN IODIDE



MIX A FEW CRYSTALS (AS MUCH AS A PEA) OF POTASSIUM IODIDE WITH ¼ TEASPOON SODIUM BISULFATE. PLACE STRIPS OF WETTED LITMUS PAPER AT THE MOUTH OF TUBE. HEAT GENTLY. IODINE IS RELEASED. ALSO HYDROGEN IODIDE—AN ACID THAT TURNS BLUE LITMUS RED.

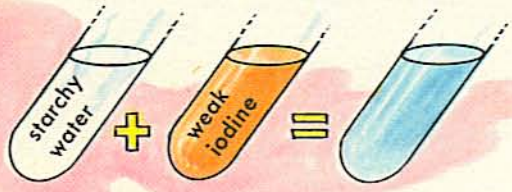
STARCH TEST FOR IODINE

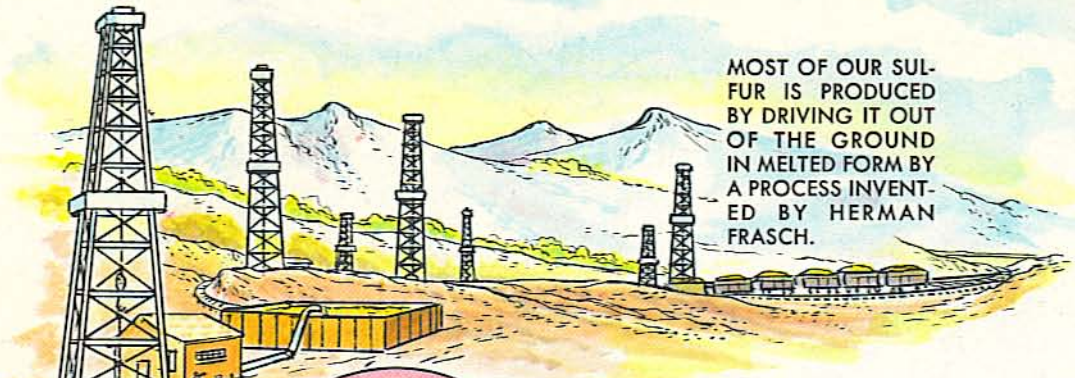


SHAKE UP A PINCH OF STARCH WITH COLD WATER IN A TEST TUBE. ADD TO HOT WATER. BRING TO A BOIL. COOL. POUR DROP OF MIXTURE INTO 10 ml WATER. ADD DROP OF IODINE SOLUTION. BRIGHT BLUE COLOR RESULTS.

REMOVING IODINE STAIN

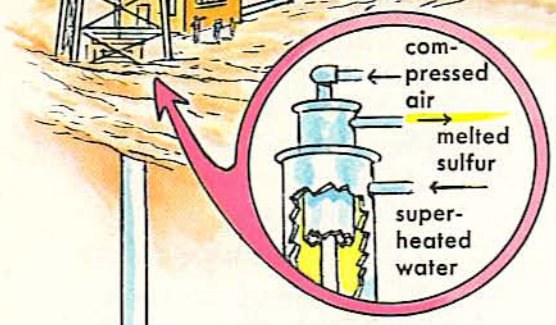
PAINT PAPER WITH IODINE. DISSOLVE A FEW CRYSTALS OF SODIUM THIOSULFATE ("HYPO") IN WATER. PAINT WITH THIS SOLUTION OVER THE BROWN COLOR. YOU WILL GET WHITE LETTERS AS HYPO FORMS COLORLESS COMPOUND WITH IODINE.





MOST OF OUR SULFUR IS PRODUCED BY DRIVING IT OUT OF THE GROUND IN MELTED FORM BY A PROCESS INVENTED BY HERMAN FRASCH.

S SULFUR Element 16. Atomic wt.: 32.066. Density: 2.07. Yellow crystals. Insoluble in water. Melts at 119°C. Boils at 444°C. Burns in air with blue flame.



Sulfur and Its Compounds

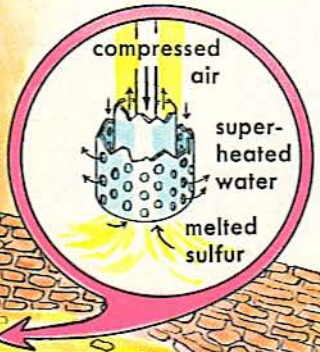
In the old days, sulfur was called "brimstone" ("burning stone" — from an old word, *brennen*, to burn). When it burned with a blue flame and a suffocating smell, people were certain that the devil himself was around.

Until fairly recently, most sulfur came from the volcanic Italian island of Sicily. But today, America produces most of the world's sulfur. About a hundred years ago, big deposits were found in Louisiana, several hundred feet underground. The problem of getting it up was solved in 1894 in a very clever way by a young German emigrant, Herman Frasch. He piped superheated water underground to melt the sulfur, then forced the melted sulfur to the top with compressed air.

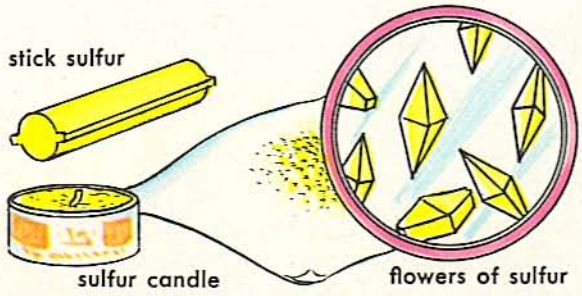
Sulfur itself is used for many purposes. By a process called "vulcanization" it turns sticky, gummy raw rubber into elastic rubber usable for automobile tires and other rubber products. Sulfur also goes into such things as matches and gunpowder and medical preparations.

But by far the greatest use of sulfur is in the preparation of sulfuric acid (H_2SO_4). This acid enters into the (CONTINUED ON PAGE 52)

SOME SULFUR USES



SUPERHEATED WATER PIPED UNDERGROUND MELTS THE SULFUR. COMPRESSED AIR FORCES SULFUR TO THE SURFACE. HERE IT IS COOLED IN LARGE WOODEN BINS.



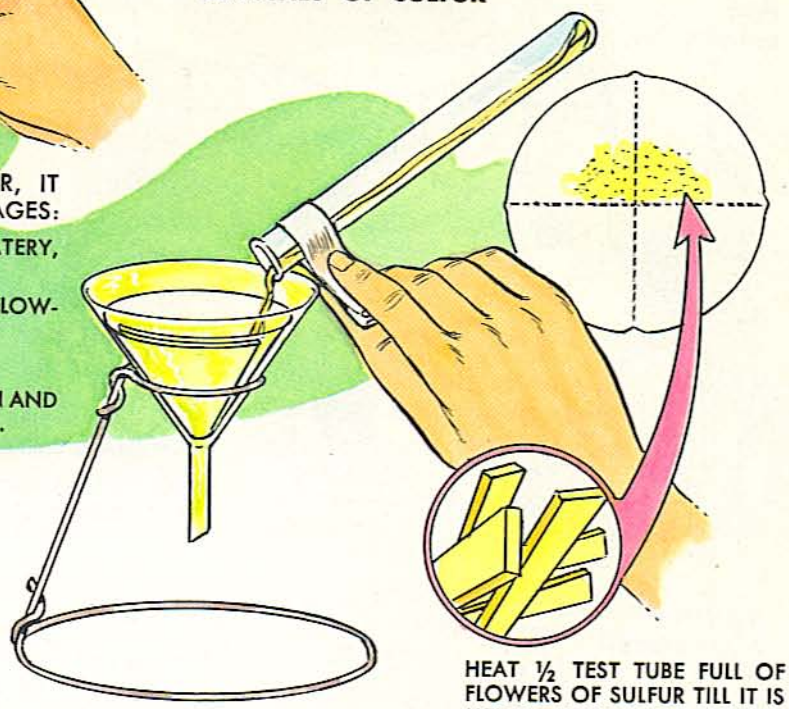
SULFUR CAN USUALLY BE BOUGHT IN THREE DIFFERENT FORMS: AS STICK SULFUR, SULFUR CANDLES, AND AS A POWDER (FLOWERS OF SULFUR). UNDER MICROSCOPE, SULFUR POWDER PROVES TO BE RHOMBIC CRYSTALS.

MELTING SULFUR

WHEN YOU MELT SULFUR, IT GOES THROUGH FOUR STAGES:

1. IT FIRST MELTS INTO A WATERY, STRAW-COLORED LIQUID.
2. IT NEXT BECOMES SLOW-FLOWING, CARAMEL-BROWN.
3. IT TURNS ALMOST SOLID.
4. IT BECOMES LIQUID AGAIN AND BOILS WITH YELLOW VAPOR.

MAKING MONOCLINIC CRYSTALS OF SULFUR



HEAT ½ TEST TUBE FULL OF FLOWERS OF SULFUR TILL IT IS MELTED WITH LIGHT COLOR. POUR MELTED SULFUR INTO A DRY FILTER. AS SOON AS CRUST FORMS ON TOP, OPEN UP FILTER PAPER. YOU WILL SEE THAT SULFUR HAS FORMED TINY NEEDLE-LIKE CRYSTALS.

CASTING WITH SULFUR

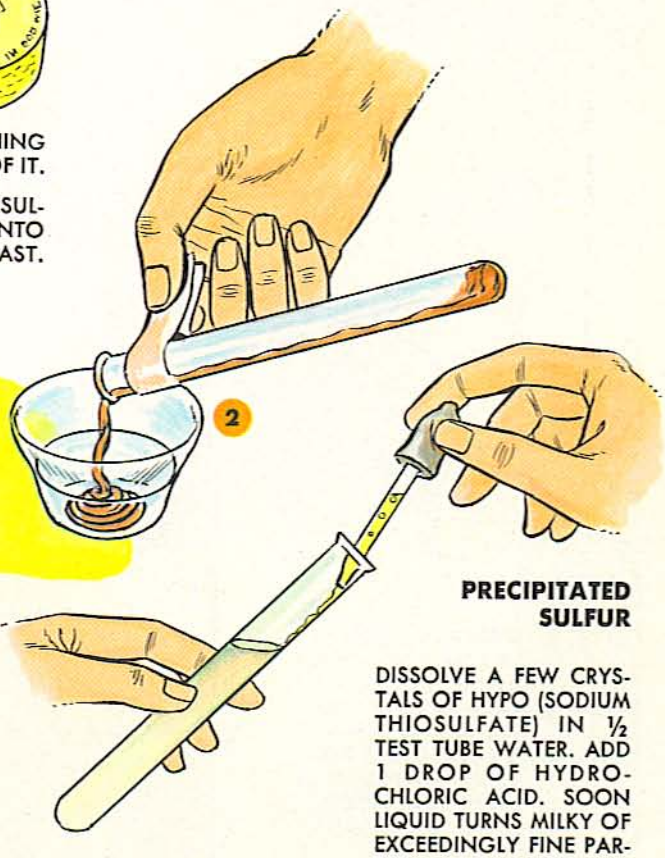
1 MAKE A MOLD FROM A NICKEL BY ATTACHING A WALL OF SCOTCH TAPE AROUND THE EDGE OF IT.

2 FILL TEST TUBE ⅓ FULL OF FLOWERS OF SULFUR. MELT GENTLY HIGH ABOVE FLAME. POUR INTO MOLD. WHEN COOLED YOU HAVE A PERFECT CAST.

PLASTIC SULFUR

1 MELT ½ TEST TUBE POWDERED SULFUR. CONTINUE HEATING. SOON IT NO LONGER FLOWS. YOU CAN TURN TUBE UPSIDE DOWN WITHOUT ANYTHING COMING OUT.

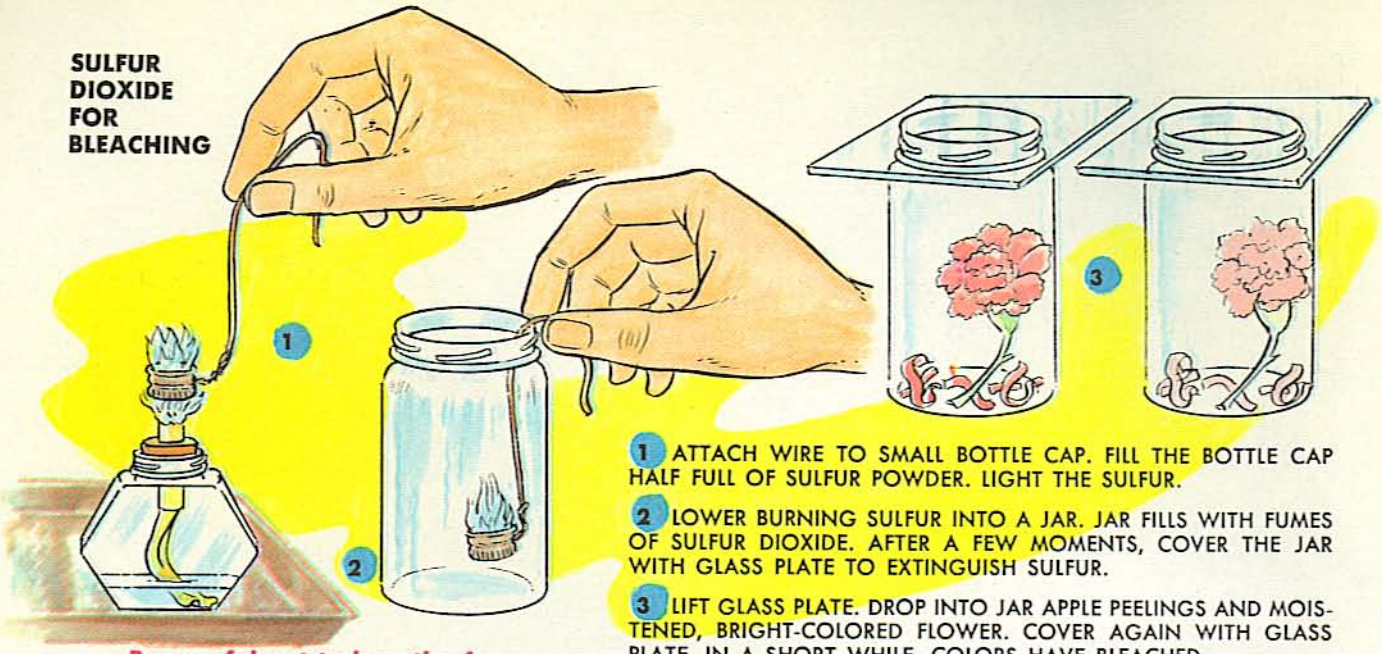
2 HEAT THE THICKENED SULFUR FURTHER UNTIL IT FLOWS FREELY AGAIN. THEN POUR THE DARK FLUID INTO COLD WATER. IT TURNS INTO A PLASTIC MASS. IN A FEW DAYS THIS AGAIN BECOMES YELLOW SULFUR.



PRECIPITATED SULFUR

DISSOLVE A FEW CRYSTALS OF HYPO (SODIUM THIOSULFATE) IN ½ TEST TUBE WATER. ADD 1 DROP OF HYDROCHLORIC ACID. SOON LIQUID TURNS MILKY OF EXCEEDINGLY FINE PARTICLES OF SULFUR.

SULFUR DIOXIDE FOR BLEACHING



1 ATTACH WIRE TO SMALL BOTTLE CAP. FILL THE BOTTLE CAP HALF FULL OF SULFUR POWDER. LIGHT THE SULFUR.

2 LOWER BURNING SULFUR INTO A JAR. JAR FILLS WITH FUMES OF SULFUR DIOXIDE. AFTER A FEW MOMENTS, COVER THE JAR WITH GLASS PLATE TO EXTINGUISH SULFUR.

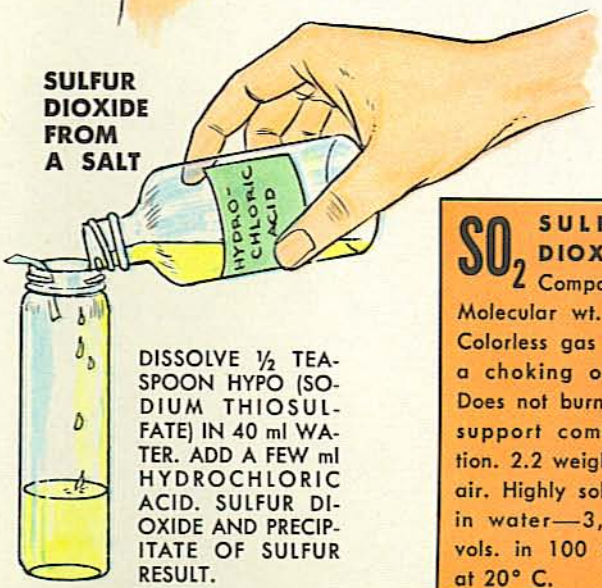
3 LIFT GLASS PLATE. DROP INTO JAR APPLE PEELINGS AND MOISTENED, BRIGHT-COLORED FLOWER. COVER AGAIN WITH GLASS PLATE. IN A SHORT WHILE, COLORS HAVE BLEACHED.

Be careful not to breathe fumes.

MAKING SULFUROUS ACID

LIGHT SULFUR IN BOTTLE CAP. LOWER BURNING SULFUR INTO JAR. WHEN JAR IS FULL OF FUMES, REMOVE SULFUR. ADD A FEW ml WATER. SHAKE. AS SO_2 DISSOLVES IN WATER IT FORMS A WEAK ACID—SULFUROUS ACID, H_2SO_3. TEST FOR ACID WITH BLUE LITMUS PAPER.

SULFUR DIOXIDE FROM A SALT



DISSOLVE ½ TEASPOON HYPO (SODIUM THIOSULFATE) IN 40 ml WATER. ADD A FEW ml HYDROCHLORIC ACID. SULFUR DIOXIDE AND PRECIPITATE OF SULFUR RESULT.

SO_2 SULFUR DIOXIDE Compound. Molecular wt. 64. Colorless gas with a choking odor. Does not burn nor support combustion. 2.2 weight of air. Highly soluble in water—3,937 vols. in 100 vols. at 20° C.

Sulfur—Continued

production — directly or indirectly — of practically every manufactured article we use today. It is used in refining gasoline, in making steel and paper, fibers and films, plastics and explosives, and thousands of other chemicals.

Sulfur Dioxide — The first step in making sulfuric acid from sulfur is to burn the sulfur.

When burning in the air, each atom of sulfur takes on two atoms of oxygen to make one molecule of sulfur dioxide gas (SO_2).

By a special, complicated process, sulfur dioxide can be forced to take on another oxygen atom and form sulfur trioxide (SO_3). With water, this makes sulfuric acid:

$$H_2O + SO_3 \rightarrow H_2SO_4$$

Hydrogen Sulfide — Many sulfur compounds have unpleasant, penetrating smells. Some of these compounds have very complex molecules — just imagine a skunk producing a chemical with this formula: $CH_3CH_2CH_2CH_2SH$! The smell of rotten eggs, on the other hand, comes from the simple compound hydrogen sulfide (H_2S).

Hydrogen sulfide is used in chemical analysis to determine what metals are found in a certain substance. It combines with metals into salts (sulfides) that can be distinguished from each other by their colors and by the way they react with acids and other chemicals.

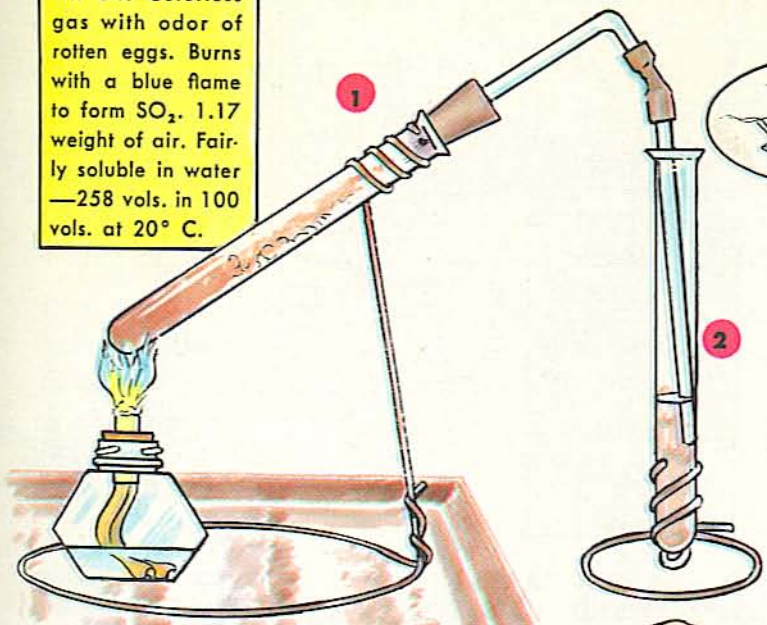
NOTE: Perform these experiments out-of-doors or before an open window. Be careful not to breathe fumes.

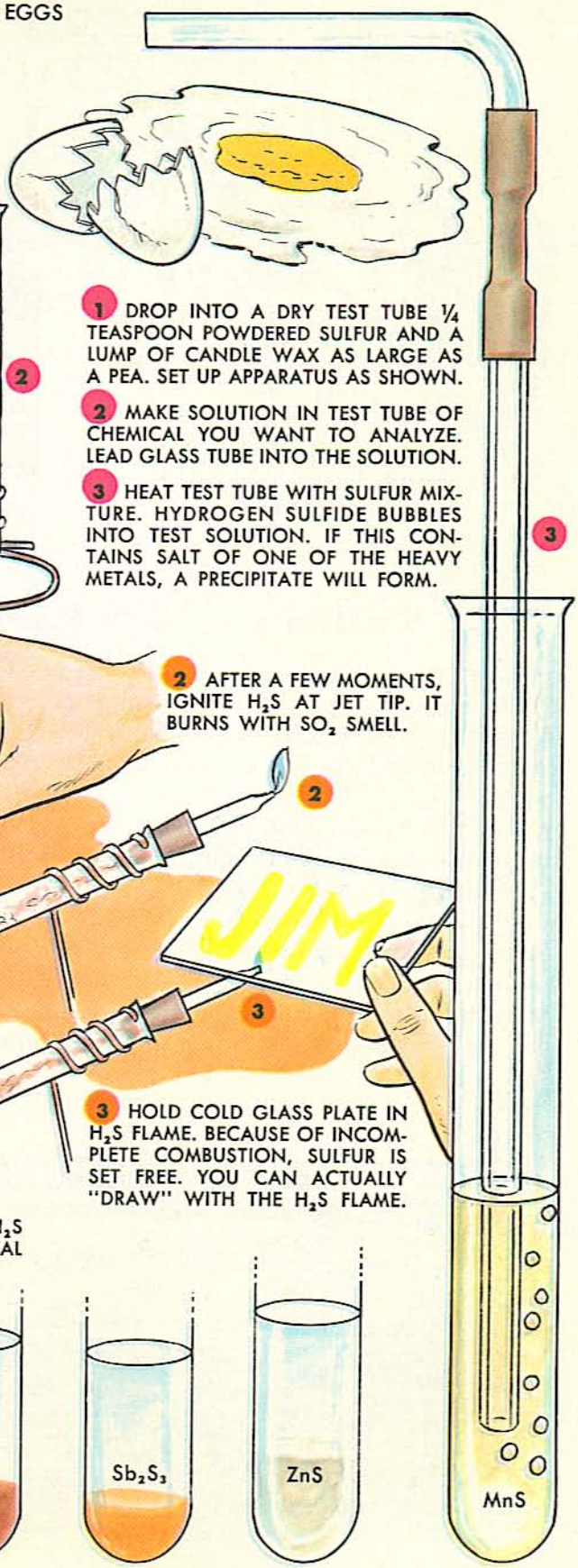
H_2S HYDROGEN SULFIDE Compound. Molecular wt. 34. Colorless gas with odor of rotten eggs. Burns with a blue flame to form SO_2. 1.17 weight of air. Fairly soluble in water —258 vols. in 100 vols. at 20° C.

HYDROGEN SULFIDE

IS AN IMPORTANT LAB TOOL FOR CHEMICAL ANALYSIS.

HYDROGEN SULFIDE HAS SMELL OF ROTTEN EGGS





1 DROP INTO A DRY TEST TUBE ¼ TEASPOON POWDERED SULFUR AND A LUMP OF CANDLE WAX AS LARGE AS A PEA. SET UP APPARATUS AS SHOWN.

2 MAKE SOLUTION IN TEST TUBE OF CHEMICAL YOU WANT TO ANALYZE. LEAD GLASS TUBE INTO THE SOLUTION.

3 HEAT TEST TUBE WITH SULFUR MIXTURE. HYDROGEN SULFIDE BUBBLES INTO TEST SOLUTION. IF THIS CONTAINS SALT OF ONE OF THE HEAVY METALS, A PRECIPITATE WILL FORM.

HYDROGEN SULFIDE FROM FeS

MAKE IRON SULFIDE AS DESCRIBED ON PAGE 22. BREAK THE TEST TUBE (IN A PAPER BAG). CRUSH THE FeS WITH A HAMMER. DROP SMALL PIECES IN ANOTHER TEST TUBE. BY ADDING HYDROCHLORIC ACID YOU MAKE H_2S.

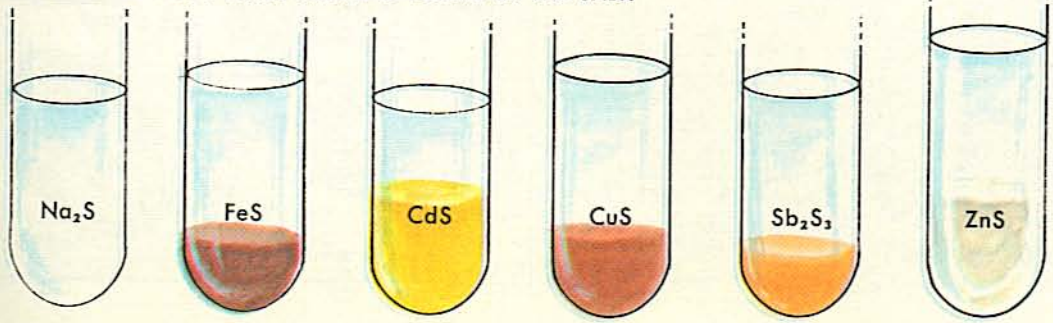
2 AFTER A FEW MOMENTS, IGNITE H_2S AT JET TIP. IT BURNS WITH SO_2 SMELL.

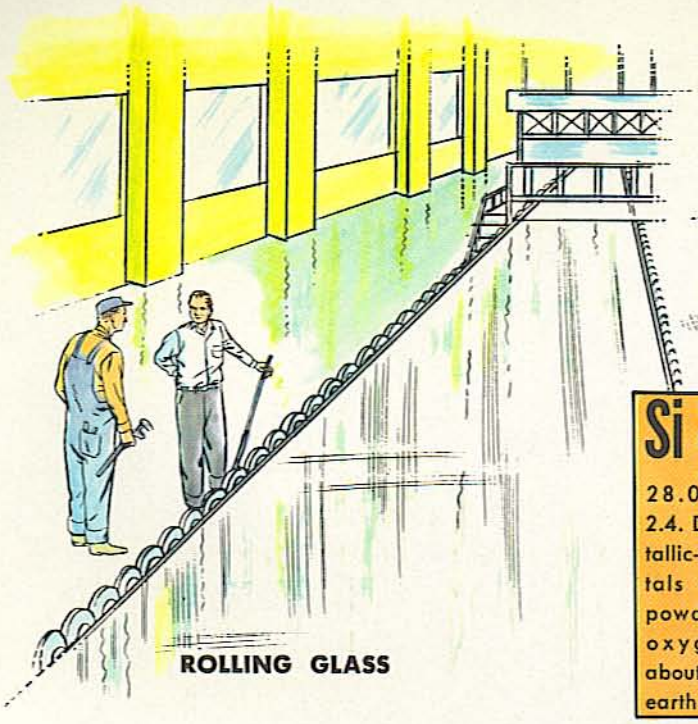
EXPERIMENTS WITH H_2S

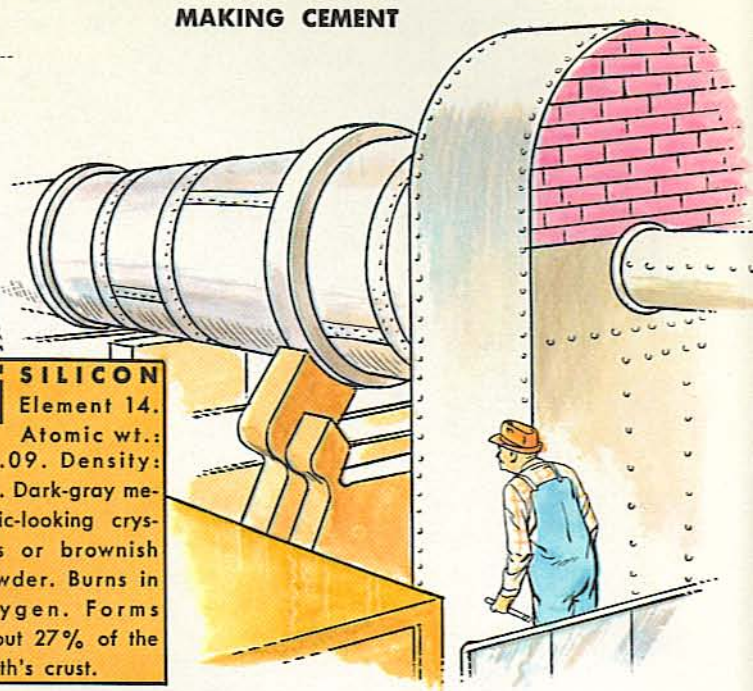
1 PUT STOPPER WITH GLASS TUBE WITH JET POINT IN TEST TUBE IN WHICH YOU MAKE H_2S. MOISTEN A SILVER COIN. HOLD IT IN H_2S STREAMING OUT OF JET. IT TURNS BLACK FROM SILVER SULFIDE.

3 HOLD COLD GLASS PLATE IN H_2S FLAME. BECAUSE OF INCOMPLETE COMBUSTION, SULFUR IS SET FREE. YOU CAN ACTUALLY "DRAW" WITH THE H_2S FLAME.

THE COLOR OF THE SULFIDE FORMED WHEN YOU LEAD H_2S INTO A SOLUTION CONTAINING A SALT OF A HEAVY METAL WILL HELP TELL YOU WHAT METAL IS FOUND IN THE SALT.







Si SILICON Element 14. Atomic wt.: 28.09. Density: 2.4. Dark-gray metallic-looking crystals or brownish powder. Burns in oxygen. Forms about 27% of the earth's crust.

Silicon—The Element You Step On

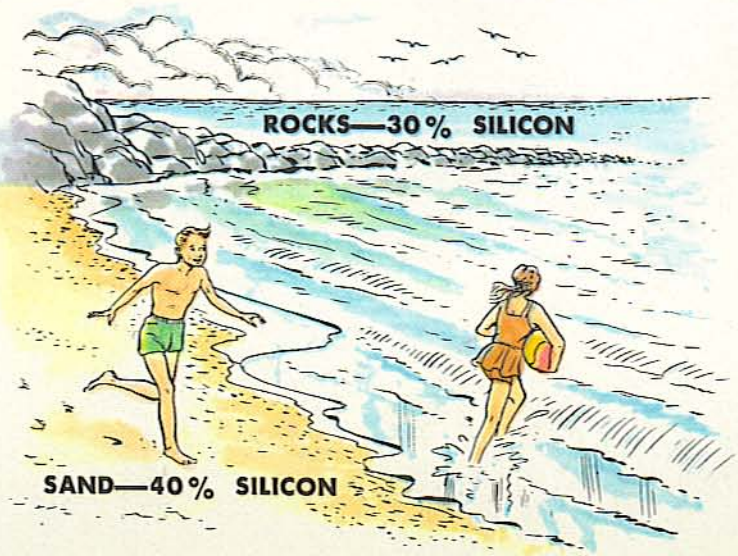
SILICON (from the Latin *silex*, flint) is the second most abundant element on earth — after oxygen. Whether you are walking on sand or clay, rock or cement, almost half of what you're stepping on is silicon.

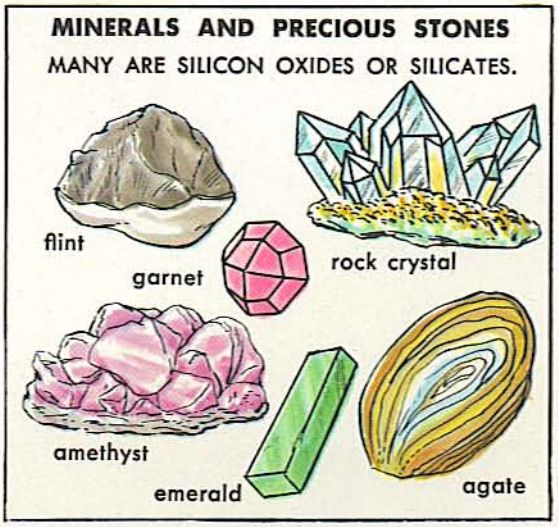
Silicon is found in nature in combination with oxygen (mostly the dioxide, SiO_2) and in different silicates (salts of various silicic acids).

With few exceptions, silicon compounds are insoluble in water. And that is a good thing for all of us. The glass of our windows and the glasses from which we drink are silicates. So are the glazes on our cups and the enamel on our bathtubs. Most glass and many glazes are made by fusing together sand (SiO_2), limestone, and soda.

The silicates of sodium and potassium dissolve in water. A concentrated solution of sodium silicate (Na_2SiO_3) is sold in hardware stores under the name of waterglass. It is used as a glue, for fireproofing wood and for preserving eggs.

Within recent years, chemists have developed a whole line of new silicon compounds called silicones. Some of them are oil-like. Some look like putty ("Silly Putty"). Still others are rubber-like. Paper and cloth can be made water-repellent by being treated with suitable silicones.

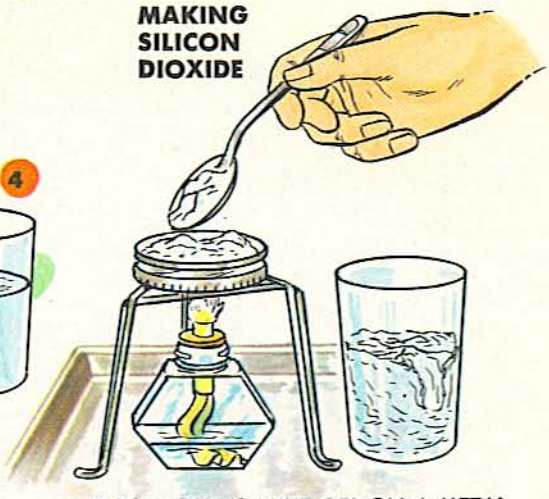




MAKING SILICIC ACID

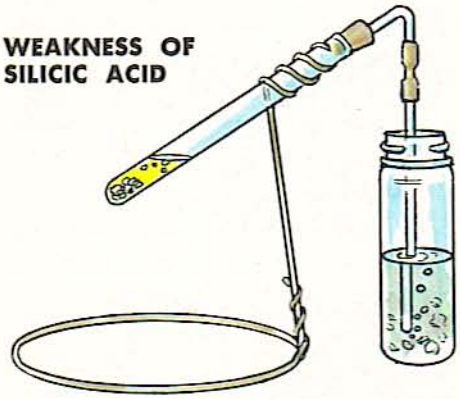
1. IN ONE GLASS, DILUTE 20 ml WATERGLASS WITH 20 ml WATER.
2. IN ANOTHER, MIX 10 ml HYDROCHLORIC ACID AND 10 ml WATER.
3. POUR THE TWO MIXTURES AT ONE TIME INTO A THIRD GLASS.
4. STAND SPOON UPRIGHT IN THE MIXTURE WHICH, ALMOST IMMEDIATELY, TURNS INTO A JELLY ("GEL") SO STIFF THAT SPOON STANDS BY ITSELF AND YOU CAN TURN THE GLASS UPSIDE DOWN.

MAKING SILICON DIOXIDE



PLACE SOME OF THE GEL ON A METAL JAR LID. HEAT. THE SILICIC ACID (H_2SiO_3) GIVES UP WATER (H_2O) AND TURNS INTO A GRAYISH-WHITE POWDER OF SILICON DIOXIDE (SiO_2).

WEAKNESS OF SILICIC ACID

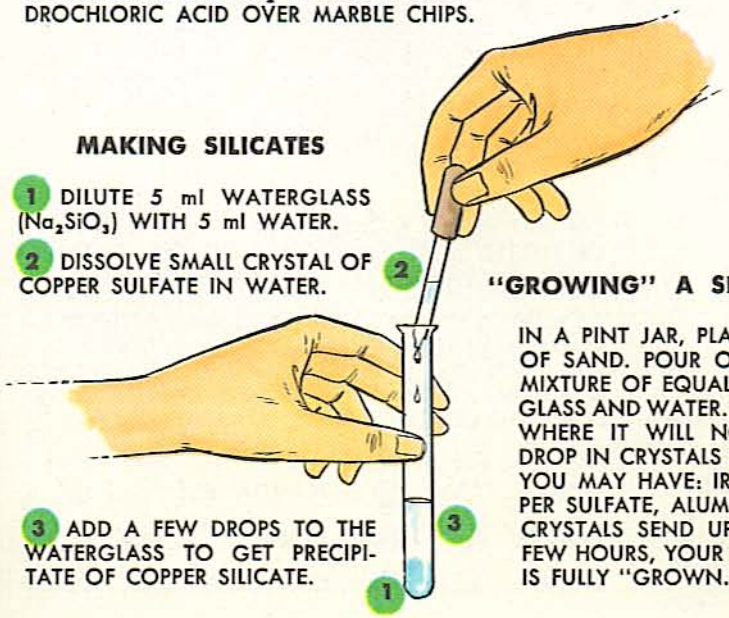


SILICIC ACID IS SO WEAK THAT CARBONIC ACID (H_2CO_3) DRIVES IT OUT OF WATERGLASS. MAKE THE CO_2 BY POURING HYDROCHLORIC ACID OVER MARBLE CHIPS.

MAKING WATER-GLASS

1. IN A TEST TUBE, MIX 1 g OF THE SILICON DIOXIDE YOU MADE, 2 g LYE (NaOH), AND 5 ml WATER. HEAT CAREFULLY, MOVING TUBE.
2. AFTER FILTERING, YOU WILL HAVE A CLEAR SOLUTION OF SODIUM SILICATE (Na_2SiO_3).

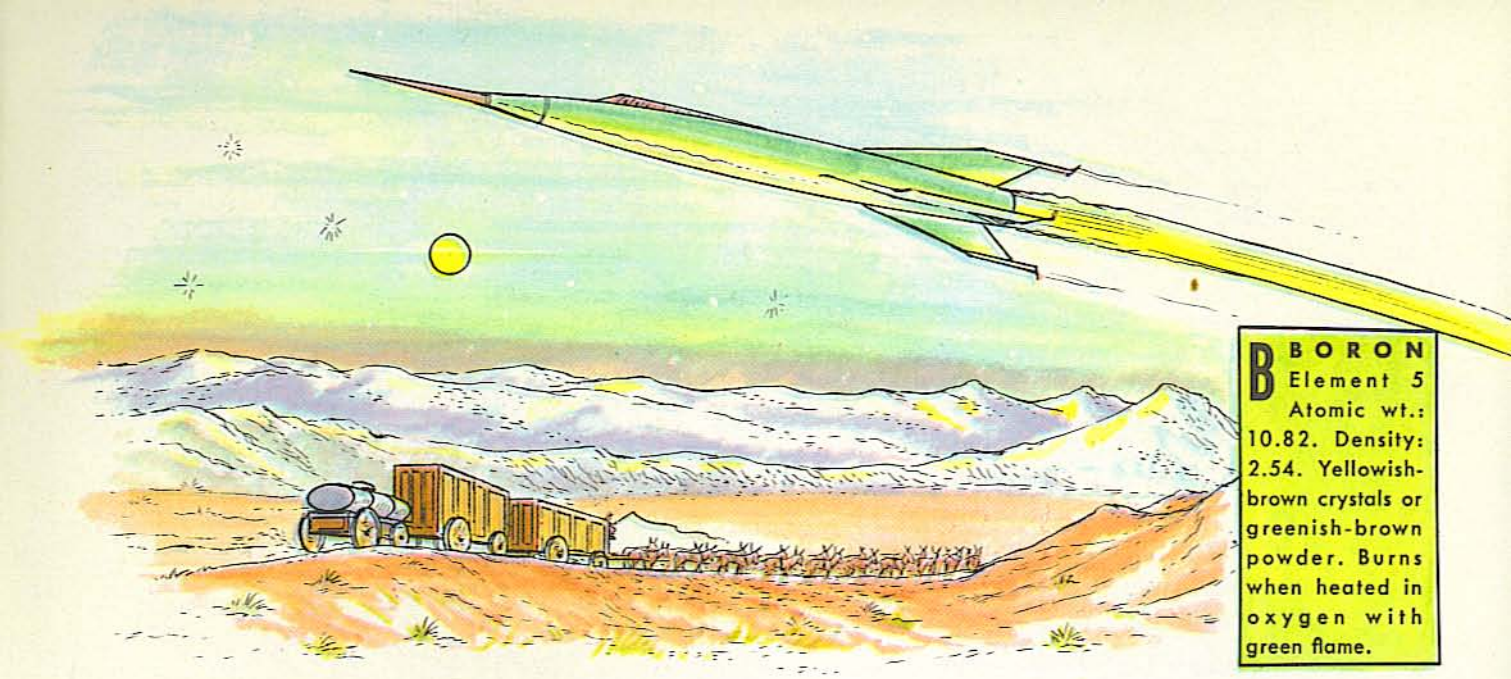
MAKING SILICATES



1. DILUTE 5 ml WATERGLASS (Na_2SiO_3) WITH 5 ml WATER.
2. DISSOLVE SMALL CRYSTAL OF COPPER SULFATE IN WATER.
3. ADD A FEW DROPS TO THE WATERGLASS TO GET PRECIPITATE OF COPPER SILICATE.

"GROWING" A SILICON "JUNGLE"

IN A PINT JAR, PLACE ½-INCH LAYER OF SAND. POUR ON TOP OF THIS A MIXTURE OF EQUAL PARTS OF WATERGLASS AND WATER. PLACE IT IN A SPOT WHERE IT WILL NOT BE DISTURBED. DROP IN CRYSTALS OF VARIOUS SALTS YOU MAY HAVE: IRON SULFATE, COPPER SULFATE, ALUM, EPSOM SALT. THE CRYSTALS SEND UP "SHOOTS." IN A FEW HOURS, YOUR SILICATE "JUNGLE" IS FULLY "GROWN."



B BORON Element 5 Atomic wt.: 10.82. Density: 2.54. Yellowish-brown crystals or greenish-brown powder. Burns when heated in oxygen with green flame.

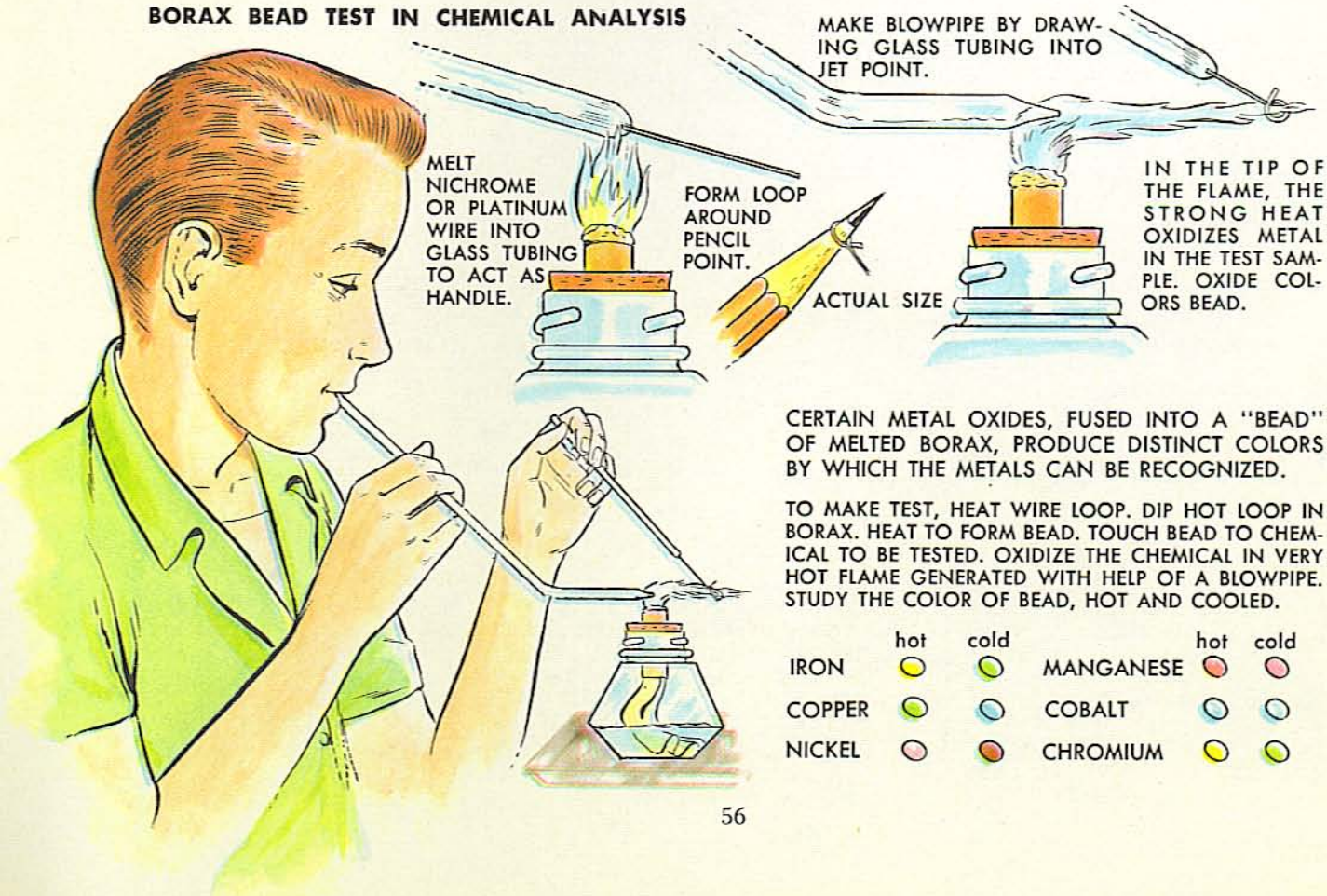
Boron—Future Rocket-Power Element?

LESS THAN A hundred years ago, a mineral called borax, containing the element boron, was carted out of Death Valley in California by twenty-mule teams — about the slowest transportation you can think of. Someday, boron may be put in zip-fuels for space missiles — the fastest form of transportation imaginable. Boron has the ability (as does carbon) to combine with hydrogen in a number of ways. When these boranes or boron hydrides burn, they develop a tremendous amount of power.

Boron can be isolated as a hard, brownish-black powder. Its carbon compound, boron carbide (B_4C), is almost as hard as diamond.

But you are probably more familiar with boron

BORAX BEAD TEST IN CHEMICAL ANALYSIS



CERTAIN METAL OXIDES, FUSED INTO A "BEAD" OF MELTED BORAX, PRODUCE DISTINCT COLORS BY WHICH THE METALS CAN BE RECOGNIZED.

TO MAKE TEST, HEAT WIRE LOOP. DIP HOT LOOP IN BORAX. HEAT TO FORM BEAD. TOUCH BEAD TO CHEMICAL TO BE TESTED. OXIDIZE THE CHEMICAL IN VERY HOT FLAME GENERATED WITH HELP OF A BLOWPIPE. STUDY THE COLOR OF BEAD, HOT AND COOLED.

	hot	cold		hot	cold
IRON			MANGANESE		
COPPER			COBALT		
NICKEL			CHROMIUM		

through two of its compounds which are found in almost every household: boric acid (H_3BO_3), used as a mild antiseptic, and borax (sodium tetraborate, $Na_2B_4O_7 \cdot 10H_2O$), used for cleaning purposes and as a water softener.

Borax has a great number of uses outside the home. It is used for soldering, for producing certain kinds of soap, and for making other boron compounds.

The glass industry uses large quantities of borax for making boron-aluminum-silicate glass. You know this kind of glass by its trade name, Pyrex. Kitchen utensils and laboratory ware made of Pyrex glass have the great advantage over ordinary glass that they can be placed directly on the fire and do not break so easily when they are subjected to sudden heating or cooling.

MAKING BORIC ACID

1 IN A CUSTARD CUP, ADD 6 g BORAX TO 15 ml WATER. BRING TO BOIL. STIR UNTIL BORAX IS DISSOLVED.

2 ADD 4 ml HYDROCHLORIC ACID TO HOT BORAX SOLUTION. STIR. REMOVE FROM FIRE. BORIC ACID CRYSTALLIZES OUT AS SOLUTION COOLS.

3 POUR CONTENTS OF CUSTARD CUP INTO A FILTER. WHEN FILTRATE HAS RUN OFF, WASH BORIC ACID REMAINING IN FILTER WITH A SMALL AMOUNT OF COLD WATER TO REMOVE NaCl ALSO FORMED IN THE PROCESS.

4 SPREAD OUT FILTER TO LET BORIC ACID CRYSTALS DRY. USE FOR EXPERIMENTS ONLY.

TESTS FOR BORIC ACID

DROP A FEW CRYSTALS OF BORIC ACID IN A CUSTARD CUP. ADD A COUPLE ml DENATURED ALCOHOL. IGNITE. STIR WITH GLASS ROD. BORIC ACID GIVES GREEN EDGES TO THE FLAMES.

YELLOW TURMERIC INDICATOR PAPER TURNS BROWN WITH BORIC ACID. YELLOW COLORING MATTER IN TABLE MUSTARD IS TURMERIC. TO MAKE TEST PAPER, DIP STRIPS OF PAPER TOWELING IN MUSTARD. WASH MUSTARD OFF. DRY STRIPS.

MAKING BORIC OXIDE

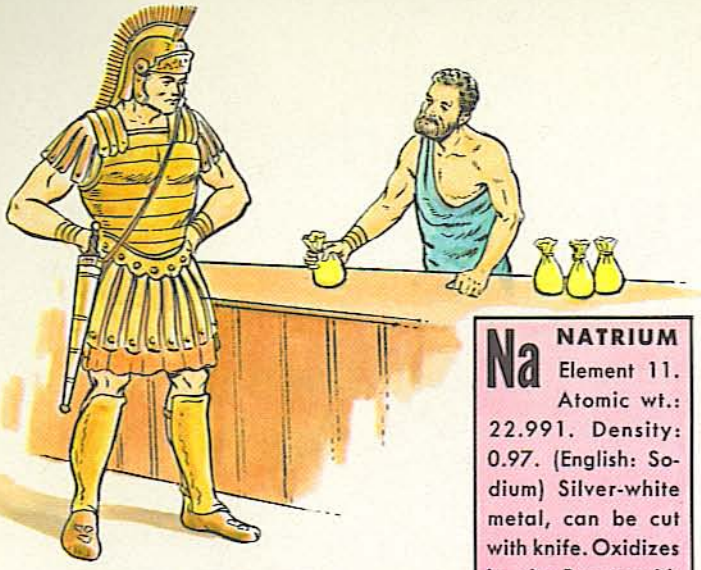
HEAT BORIC ACID IN AN OLD TEASPOON. IT GIVES OFF WATER AND TURNS INTO SYRUPY MASS OF MELTED BORIC OXIDE (B_2O_3) WHICH YOU CAN DRAW INTO THREADS WITH GLASS ROD.

Sodium and Potassium

THE SALTS of sodium and potassium have been used for thousands of years in making soap and glass and for a great number of other purposes.

Sodium chloride (NaCl) is the most common sodium salt — it is the chemical that makes ocean water "salty." Plants growing in the ocean take up so much of the sodium that people along the seacoasts of the world used to burn dried seaweed to secure "soda ash" (sodium carbonate, Na_2CO_3). Inland plants, on the other hand, pick up potassium from the soil. Inland people boiled out wood ashes in large pots to get "potash" (potassium carbonate, K_2CO_3).

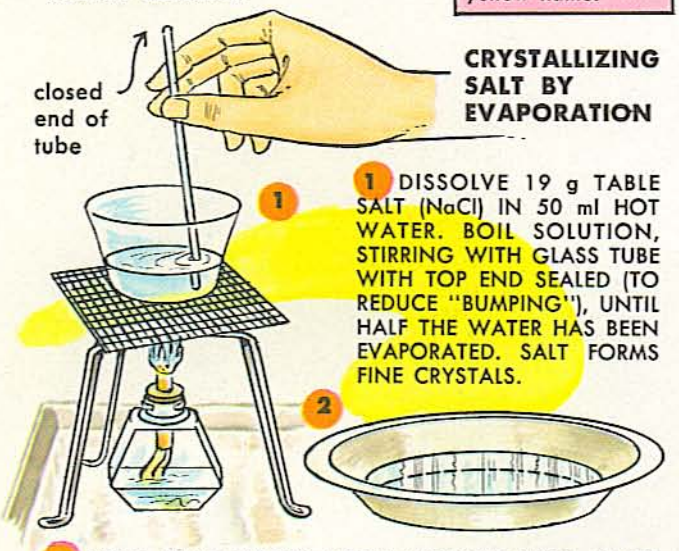
In 1807, the British scientist, Humphry Davy, succeeded in isolating the metals found in these salts. They proved to be wax-soft and silvery. He called them sodium (from soda ash) and potassium (from potash). These are still their English names. But in chemical formulas they are referred to as natrium (Na) and kalium (K) — from abbreviations of the Arabic names of the ashes: *natrun* and *al qili* (alkali).



"SALARY" COMES FROM *SALARIUM* —THE WAGES PAID IN SALT TO ROMAN SOLDIERS.

Na **NATRIUM** Element 11. Atomic wt.: 22.991. Density: 0.97. (English: Sodium) Silver-white metal, can be cut with knife. Oxidizes in air. Reacts with water. Burns with yellow flame.

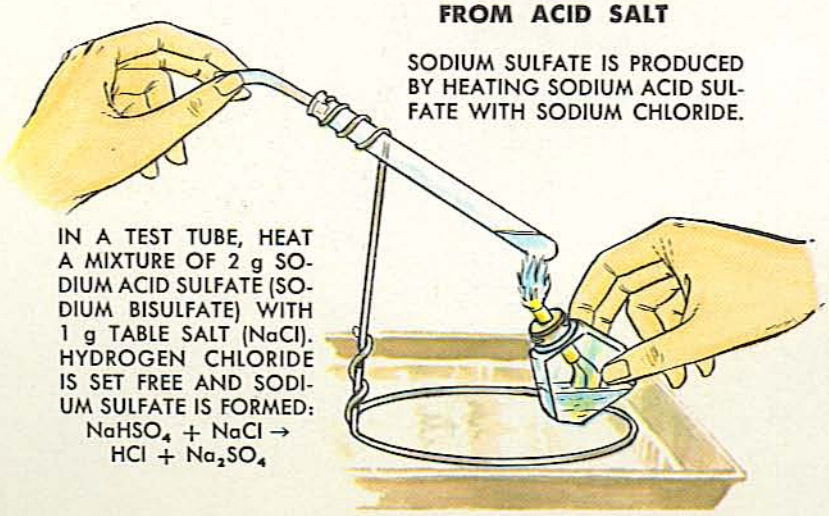
CRYSTALLIZING SALT BY EVAPORATION



1 DISSOLVE 19 g TABLE SALT (NaCl) IN 50 ml HOT WATER. BOIL SOLUTION, STIRRING WITH GLASS TUBE WITH TOP END SEALED (TO REDUCE "BUMPING"), UNTIL HALF THE WATER HAS BEEN EVAPORATED. SALT FORMS FINE CRYSTALS.

2 POUR CLEAR LIQUID INTO LARGE PIE PLATE. PLACE IN SUNNY WINDOW FOR WATER TO EVAPORATE SLOWLY. THE CRYSTALS FORMED WILL BE MUCH LARGER.

MAKING NORMAL SALT FROM ACID SALT

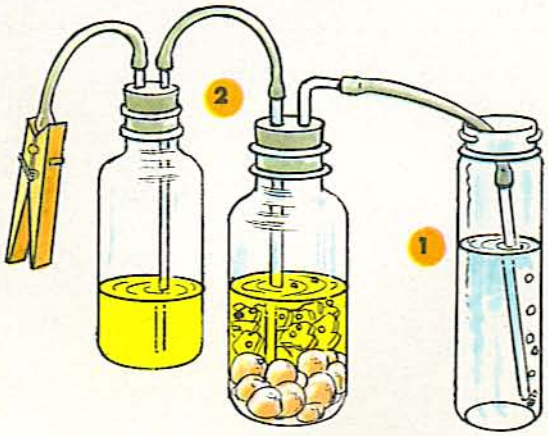


SODIUM SULFATE IS PRODUCED BY HEATING SODIUM ACID SULFATE WITH SODIUM CHLORIDE.

IN A TEST TUBE, HEAT A MIXTURE OF 2 g SODIUM ACID SULFATE (SODIUM BISULFATE) WITH 1 g TABLE SALT (NaCl). HYDROGEN CHLORIDE IS SET FREE AND SODIUM SULFATE IS FORMED:

$$NaHSO_4 + NaCl \rightarrow HCl + Na_2SO_4$$

MAKING ACID SALT FROM NORMAL SALT



THE ACID CARBONATE ($NaHCO_3$) IS MADE BY LEADING CO_2 TO NORMAL CARBONATE (Na_2CO_3).

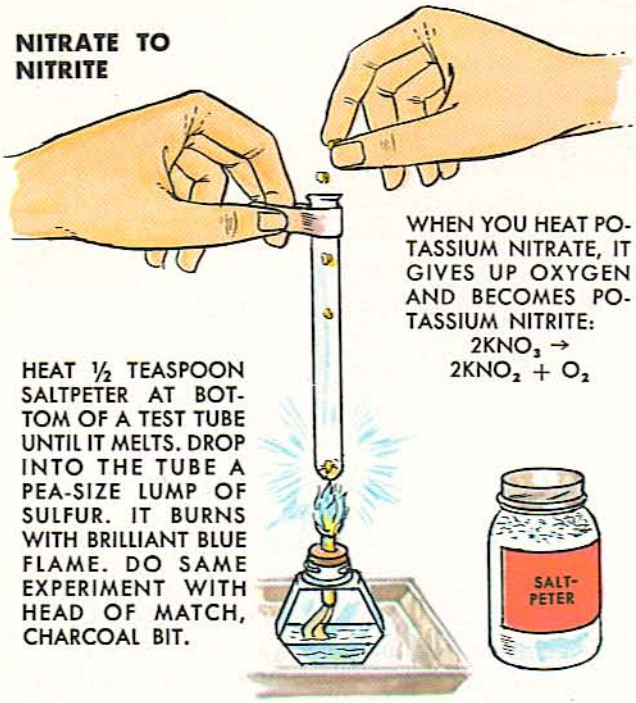
1 MAKE SATURATED SOLUTION BY SHAKING 3 TEASPOONS WASHING SODA IN 30 ml COOL, BUT NOT COLD, WATER. FILTER IT.

2 SET UP APPARATUS FOR MAKING CO_2 AS SHOWN ON PAGE 31. LEAD CO_2 INTO SODA SOLUTION FOR 10 MINUTES. THEN SET ASIDE. SHORTLY $NaHCO_3$ CRYSTALS APPEAR.

K KALIUM Element 19. Atomic wt.: 39.100. Density: 0.87. (English: Potassium) Silver-white metal, so soft it can be cut with knife. Oxidizes in air. Reacts with water. Burns with violet flame.

PIONEER WOMEN MADE POTASH FROM WOOD ASHES.

NITRATE TO NITRITE



WHEN YOU HEAT POTASSIUM NITRATE, IT GIVES UP OXYGEN AND BECOMES POTASSIUM NITRITE:

$$2KNO_3 \rightarrow 2KNO_2 + O_2$$

HEAT ½ TEASPOON SALTPETER AT BOTTOM OF A TEST TUBE UNTIL IT MELTS. DROP INTO THE TUBE A PEA-SIZE LUMP OF SULFUR. IT BURNS WITH BRILLIANT BLUE FLAME. DO SAME EXPERIMENT WITH HEAD OF MATCH, CHARCOAL BIT.

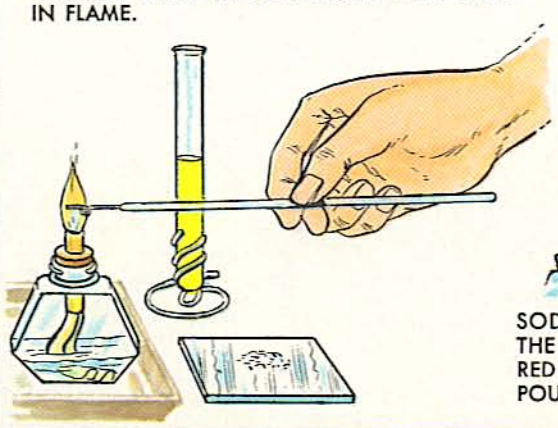
MAKING POTASH

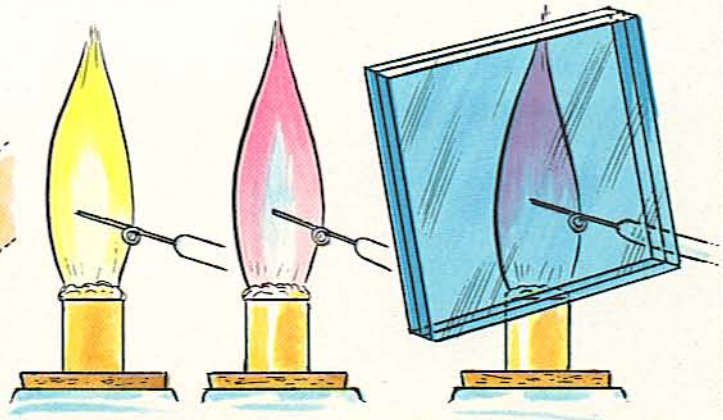
1 STIR UP SEVERAL TEASPOONS OF FRESH WOOD ASHES WITH WARM WATER. SKIM OFF WOOD REMAINS.

2 FILTER THE MIXTURE OF ASHES AND WATER. COLLECT FILTRATE IN A CUSTARD CUP. EVAPORATE MOST OF WATER. THEN COOL TO PERMIT K_2CO_3 TO CRYSTALLIZE OUT.

FLAME COLOR TEST FOR SODIUM AND POTASSIUM

THE COMPOUNDS OF CERTAIN METALS GIVE DISTINCT COLORS TO A FLAME. DIP NICHROME WIRE IN HCl TO CLEAN IT. HEAT IT. THEN DIP LOOP IN COMPOUND AND HOLD IN FLAME.





SODIUM COMPOUNDS GIVE THE FLAME A BRIGHT YELLOW-RED COLOR. POTASSIUM COMPOUNDS GIVE VIOLET FLAME.

TO SEE VIOLET COLOR OF POTASSIUM IN MIXTURE WITH Na, USE BLUE GLASS TO SCREEN OUT YELLOW OF Na.