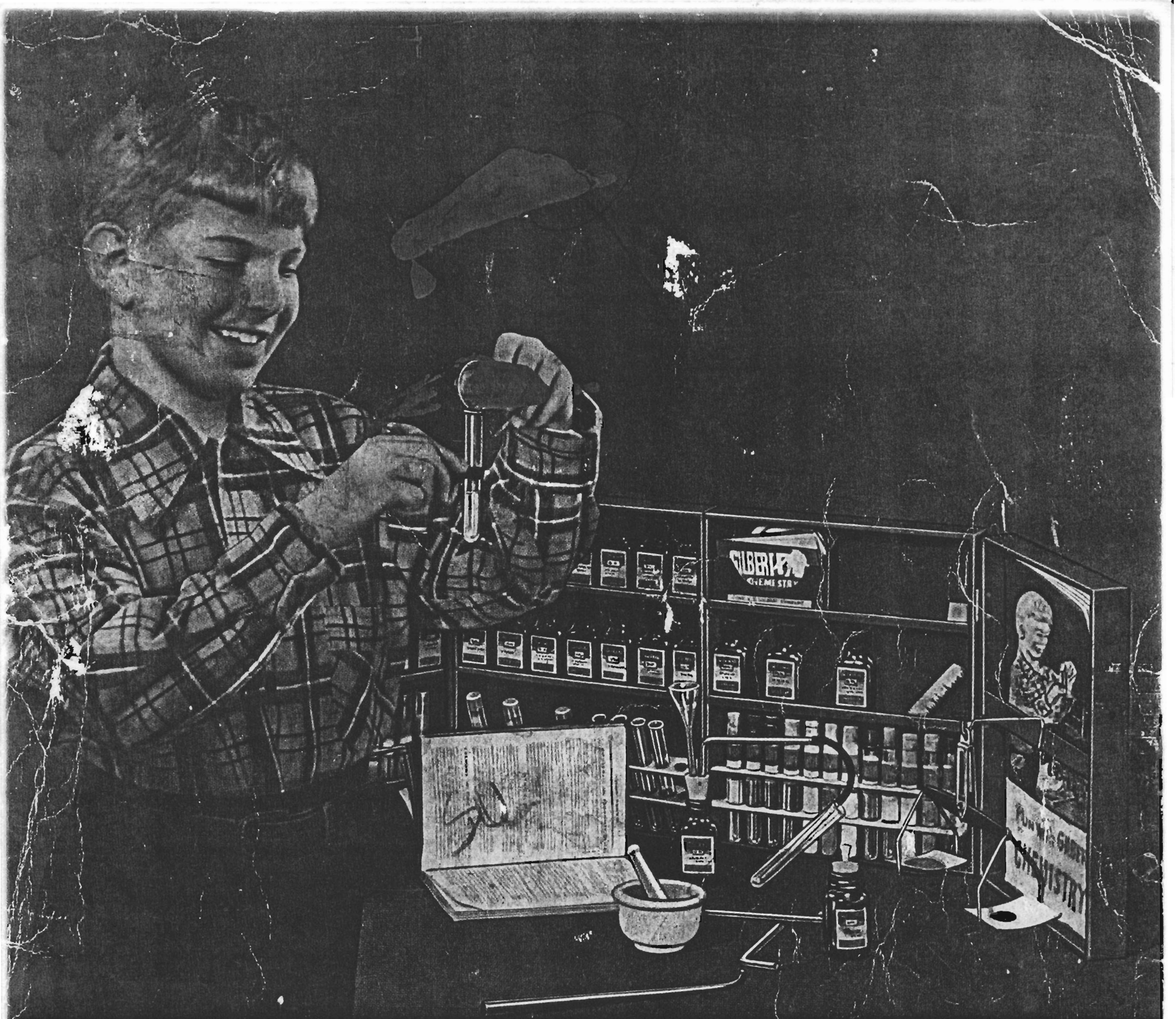

Fun with Gilbert CHEMISTRY

Fun with Gilbert CHEMISTRY

by

Treat B. Johnson

Yale University, Ph.D. 1901

in collaboration with

Alfred C. Gilbert

Yale University, M.D. 1909

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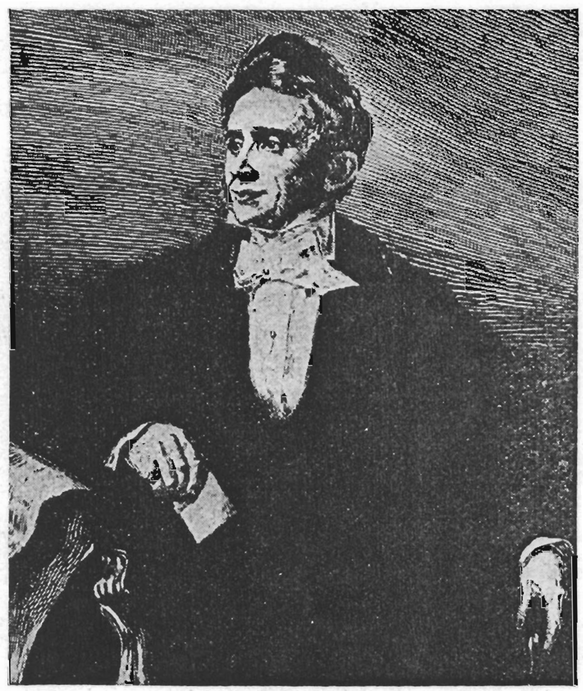
Thomas A. Edison

Thomas A. Edison (1847-1931) when only 12 years of age was working as a railroad newsboy and studying and experimenting in the fields of chemistry and physics in his spare time. At the time of his death, he had taken out well over a thousand patents and has given the world some of the outstanding scientific inventions of the twentieth century including the electric light, the storage battery, electrified railroads, the phonograph, moving pictures and countless other now common devices.

Charles Goodyear

Charles Goodyear (1800-1860) as a young man was very much interested in the chemistry of rubber and was always experimenting in an attempt to discover a method of treatment by which rubber could be made into articles that would stand extremes of hot and cold. One day he accidentally dropped some india rubber and sulphur on a hot stove and thereby discovered the process for the vulcanization of rubber.



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CONTENTS

CAUTION

Gilbert Chemistry Sets are intended only for children who can read and understand this Instruction Book.

The Gilbert Chemistry Set is constructed and organized to serve as a method for the instruction and entertainment of boys and girls in their homes. Every effort has been made to exclude the use of dangerous and poisonous chemicals and to place in the hands of our boys and girls only harmless material. This does not mean that any chemical incorporated in the set can be tasted or swallowed without injury to the body. The authors suggest that all experimentation be carried out cautiously and according to directions, especially when an operation such as heating is involved, or when gases are produced by the reactions.

Before performing experiments in the home be sure to spread a thick layer of fireproof protective material, such as a piece of linoleum, over the table, so that hot liquids, candle grease etc. will not injure the table.

Always read the directions of an experiment entirely through before starting to perform it. By following this rule many mistakes will be avoided.

When heating a test tube never point the open end at yourself or anyone nearby, as it may suddenly boil over, causing burns or injuring clothing. Never smell at the open end of a test tube while heating, or have your face near it.

TO THE PARENTS

The manufacturer of this set designed it for boys and girls from seventh and eighth grade to high school age who are qualified to read and understand the descriptive matter and the specific directions for performing the many experiments. Boys and girls of this school age are old enough to understand and observe the common rules of caution.

Parents are urged to supervise and direct the experimentation of their children. Many experiments may result in failure unless a good technique is developed and it is, therefore, urgent that parents supervise the first experiments undertaken. Carelessness on the part of boys and girls should be discouraged and avoided.

Treat B. Johnson
Alfred C. Gilbert

FOREWORD

Professor Treat B. Johnson

Prof. Johnson has been associated with the Chemistry Department of Yale University since 1898. He received the degree of Ph.B. from the Sheffield Scientific School of Yale in 1898, and the Ph.D. degree from The Yale Graduate School in 1901. Since this later date he has been continually engaged at Yale in the teaching of organic chemistry and in the promotion and development of advanced research in the fields of Organic Chemistry and Biochemistry.

Prof. Johnson is a member of the National Academy of Sciences, The National Research Council, The American Chemical Society, The American Institute of Chemists, The Connecticut Academy of Arts and Sciences, The American Society of Biological Chemists, Sigma Xi and Alpha Chi Sigma.

He now occupies the academic rank in Yale University of Sterling Professor of Chemistry Emeritus having been retired from active service in July 1943.

Fun With Chemistry

The object in writing this elementary book is to present the simpler basic principles of the Science of Chemistry in a form which will appeal to boys and girls, and arouse in them a desire to acquire an understanding and appreciation of some of the fundamental laws of nature. The book has necessarily been arranged and written in simple style to stimulate the boy's and girl's interest and curiosity without creating the feeling, on the part of the reader, that he is undertaking a laborious problem without any pleasure. It is earnestly hoped that this set will enable any intelligent boy or girl of preparatory school age to have a lot of fun in doing the various experiments and at the same time learn something of the Science of Chemistry.

Nature's Wonderland

Nature has a wonderful and interesting story to tell. It is important for us to become acquainted with her physical laws. Nature never reveals any of her secrets and fundamental laws or mechanisms to man's curiosity. Whatever man has learned to date has been the result of deep, searching investigation on the part of persons of extraordinary and original powers of observation. As a result, nature's hidden secrets are uncovered by only a few persons. It seems to be true that in mass we cannot be left to ourselves to develop normally. We need leaders to guide us. If none are born into this world who will acquire the knowledge to instruct and lead man rightly he cannot make progress. He will then be influenced by ulterior motives, and be guided by those who contrive to lead him to ruin.

If we are to formulate an educational policy, based on the above philosophy, it is quite evident that every boy and girl should be guided and stimulated to play a part in life. Different programs can be drawn up to meet individual requirements. It is not to be expected that all boys and girls can attain the same goal. Some are far more capable than others, and it is the duty of their parents and teachers to help youth in making sound decisions. If you want to get more and make a success of life there is one basic fact to be faced namely: "You know want you want; you can go practically anywhere, do almost anything you wish, but to make a success of your venture you must know what you want before you start. Life will give you what you want if you ask for it."

These principles apply to the development of every normal boy and girl in every circumstance. Make up your mind what you want. It is your right and duty. Boys and girls are put into this world to do something and—be something. Let us teach them in their early life not to be held back by their own indecisions.

16TH CENTURY CHEMISTRY LABORATORY

Note water bath in center, hooded stills at lower right showing one in action. Boy in foreground is operating a pestle suspended from an elastic pole. Man in right background is working a press. At left is a still and filter stand.

Part One

CHEMISTRY AS A SCIENCE

THE first literary work in which the word "Chemistry" is found was written by Plutarch, a Roman historian who lived from 46-120 A.D. In a treatise entitled "Isis or Osiris"—that philosopher mentions that "Egypt" in the dialect of the country, was called the same name as the black of the eye, "Chemia," and from this he infers that the word means "Black" in the Egyptian language. Some science historians believe that our word "Chemistry" means "The Egyptian Art." Others think that the word was coined to mean "The Black Art." Still others think that the word meant "The Dark or Hidden Art." Another school of thinkers believes that the word has no connection with Egypt at all, but that it comes from the Hebrew word "Chaman," meaning mystery. Another possible derivation, according to some historians is from the Arabic word "Chema" meaning to hide, hence "The Hidden Science." In fact, a book of secrets was written in the time of the ancient Arabians called "Kemi." Probably no one will ever know definitely which one of these possible derivations is the correct one.

Origin of Chemistry as a Science

According to some historians, the origin of chemistry as a science dates back to the time of Tubal Cain, the father of workers in metal. Credit is also given to Hermas, the Egyptian god of art and sciences. His son is said to have colonized Egypt, which was foremost in the knowledge of chemistry in those ancient days for they had developed the arts of making glass, pottery, colors, embalming fluids and other practical products to a high degree, and the early Egyptians can really be said, therefore, to have had an advanced knowledge of applied chemistry. Then Paracelsus, the Greek Physician, carried the study along and discovered the influence of chemistry upon medicine in the treatment of human ills, and it was through him that the action of several inorganic salts upon the human system became known. Following this period a long time elapsed, hundreds of years, during which time contributions were occasionally made by unknown workers in science, but which really had little influence upon the development of modern chemistry.

It was not until the early part of the eighteenth century that the scientists of the central European countries and the English Empire began to contribute fundamental knowledge, which laid the foundation and paved the way for the development of this wonderful science. The Frenchman, Lavoisier (1743-94) may really be credited with being the father of modern chemistry. His first chemistry work was on "Analysis of Gypsum" done in 1765. During the Reign of Terror in 1794 he was put to death on the guillotine.

There is hardly a science today that has greater economic influence, or holds more fascinating interest to scientists throughout the world than chemistry. If we are to unravel the secrets of our wonderful world and life, there is no science that will enable us to understand and correctly interpret these hidden things of nature that most of us think are magical and mysterious, like a knowledge of chemistry.

Present Day Chemistry

No large and progressive manufacturing industry can cope with its competitors today without a trained chemist to advise and assist in its development and to analyze the raw materials which it buys. The present-day physician without a knowledge of chemistry would be incompetent and unable to maintain an acceptable professional standing as a practitioner of medicine.

The great industrial problems involved in the manufacture of synthetic drugs, dyes, perfumes, explosives, essential oils, and soil fertilization chemicals are a few examples of the every-day problems of the modern research laboratories in applied chemistry. The regulation of our food supply calls for the services of thousands of experienced technicians who are employed as chemists by industry, municipalities, state and national governments. If we would have our country improve its present standard of living and at the same time accommodate itself to an increasing population, we must maintain, on an even more liberal scale than ever before, great laboratories of science devoted to the study of chemistry. The trained men and women working in these laboratories are among our priceless possessions.

A Career in Chemistry

Probably most boys are interested in science because they just naturally think they will like science. This is a perfectly good and sufficient reason in itself. At the same time, a boy of intelligence who becomes interested in science would like to be reassured, no doubt, that science offers a really important field for service in the interests of human welfare. From the far-sighted point of view, the public is much better off today than it was before science was developed, and so it always will be. Every boy and girl should be impressed with this fact and be made to realize that science creates jobs, and that its application makes life more comfortable and more interesting.

A Chemist is looked upon today as a professional man to be treated with respect, and there is a growing desire among all people to know more about this science.

To satisfy the youthful thirst for chemical knowledge, and to afford the pleasure to be derived from the intelligent performance of simple experiments, is one of the aims of this manual. The second aim is to develop the power of scientific reasoning and to give to the boy an elementary knowledge of the fundamental principles upon which modern chemistry is based.

Rules of the Game

The experiments must be carried out with accuracy in order to obtain satisfactory results. Remember that nature is exacting in her methods of operation, and it is the problem of the scientist to seek the truth and operate according to the "rules of the game" so to speak, by careful experimentation. The author would, therefore, urge that you think out for yourself, when you are performing experiments, first the purpose of the experiment, second, consider carefully the results obtained, and third, draw some conclusions as to what the results really mean to you. By so doing you will develop your imagination and an investigative mind. The performance of your experiments will provide fun for you, at the same time contribute to your knowledge, and also advance and develop the science of chemistry. The young experimentalists of today will be the productive scientists of the future.

DEFINITIONS OF SOME COMMON CHEMICAL TERMS

Acids

Chemical compounds which have a sour taste. They turn blue litmus red. They unite with bases to form salts. They all contain hydrogen, which is a gas.

Atom

The smallest particle of an element which enters into chemical combination. Atoms are extraordinarily small. We can never hope to see one, even with powerful microscopes.

Atomic Weight

Relative weight of an atom compared with an atom of hydrogen as a standard. Since hydrogen is the lightest known element, the weight of its atom is taken as 1. When we say that the atomic weight of oxygen is 16 we mean that the atoms of oxygen are sixteen times heavier than those of hydrogen.

Base

A base is water in which one-half of the hydrogen has been replaced by a metal. For example, water is HOH. A base, sodium hydroxide is Na OH. Bases are also known as alkalies. They combine with acids to form a salt of the metal and water. Bases will turn red litmus paper blue.

Chemistry

The science of chemistry has for its object the accurate investigation of all changes in the identity of substances and the laws, causes and effects of such changes.

Chemical Change

A change which destroys the identity of the substance or substances acted upon.

Chemical Compound

A union of two or more substances in definite quantities, combined so as to form a new and distinct substance which is unlike either of the substances which formed it.

Chemical Equation

When chemical substances react upon or unite with one another, definite transformations take place which can be expressed in the form of a chemical equation. Thus we may express the action of hydrochloric acid on calcium carbonate to form calcium chloride, water and carbon dioxide as follows:

$$2\ HCl + CaCO_3 = CaCl_2 + H_2O + CO_2$$

An equation is an abbreviated form of what takes place in a chemical reaction.

Chemical Affinity

Property which elements have for uniting with one another.

Decolorize

To bleach or whiten—to remove the color from a liquid or solid.

Decompose

The process by which a compound breaks up into simpler parts—usually through the action of heat.

Deodorize

To remove an odor or smell—especially the odor which results from impurities.

Dissociate

The process by which a compound breaks up into ions when dissolved in water.

Electrolysis

The decomposition or breaking up of a chemical compound by means of an electric current.

Element

A substance which cannot be separated into simpler parts.

Evaporate

To change a liquid or solid into a vapor or gas. This is usually done with heat. Minerals, salts or ash often remain behind. Many liquids will evaporate on simple exposure to the air.

Immerse

To dip or plunge into anything that surrounds or covers—especially a liquid.

Ion

An atom or group of atoms which carries a certain amount or charge of electricity.

Law of Conservation of Matter

Matter can neither be created nor destroyed. For example, if we burn a piece of coal, the weight of ashes and gases formed after burning is exactly equal to the weight of the coal before burning, plus the oxygen consumed in the combustion process.

Law of Definite Proportions

Atoms unite with one another in definite though frequently in two or more different proportions. For example, carbon, sulphur and arsenic form two distinct oxides, CO and CO SO_2 and SO_3, and As_2 SO_3 and As_2 SO_5.

Mixture

A mass of two or more ingredients, the particles of which are separate, independent and uncompounded with each other, no matter how thoroughly and finely they are mixed. There is no chemical union as there is in a compound.

Molecule

The smallest particle of a chemical compound which is capable of existence.

Physical Change

A change which does not affect the identity of the substance or substances acted upon.

Precipitate

An insoluble substance separated from a solution by the action of some substance which is added to the solution. The precipitate may fall to the bottom (hence the name which means "to throw down"), or it may float in the liquid.

Salts

Compounds formed by the combination of acids and bases.

Solution

The process by which a body, whether solid, liquid or gaseous, is absorbed in a liquid and diffused or spread throughout the liquid. The liquid is called the solvent.

Symbol

For convenience, elements are designated by symbols. Each symbol stands for one atom of an element: as S for sulphur, Pb for lead (latin "plumbum"). NaCl is the chemical formula of sodium chloride or common table salt.

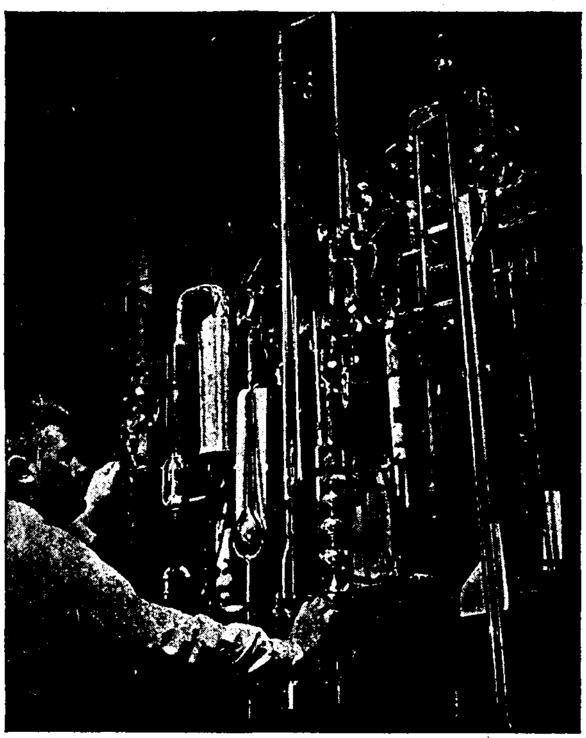
CHEMISTRY CLASSIFICATION

Inorganic Chemistry

Metals, minerals, ores, etc., which are found in nature occurring in the crust of the earth, are classified as inorganic materials. They are not combustible in the sense that they can be burned like carbon to gaseous products. They represent a group which was originally spoken of as mineral substances and are distinguished from those products or substances which originate directly or indirectly from living organisms. Iron, copper, glass and the ore pyrite, for example, are typical inorganic substances. All materials of this nature are treated under a specific classification which we designate—*inorganic chemistry.*

Organic Chemistry

It was the French chemist, Lavoisier (1743-94) who showed that in spite of their great number, nearly all vegetable products occurring in nature are composed of three elements—carbon, hydrogen, and oxygen—whereas animal substances, which also consist, for the most part, of these same three elements, contain nitrogen, and in some cases phosphorus, sulphur and iodine. All such products were shown to be not only peculiar in their composition, but also were combustible. This discovery of Lavoisier and later workers led to the belief that all animal and vegetable substances in nature were produced under the influence of a *vital force* and that their formation in nature was regulated by laws which were different from those which governed the formation of mineral substances. For this reason, therefore, compounds obtained from animals and plants, either directly or indirectly, were called organic compounds and a study of the products of this type was classified under the designation "Organic Chemistry." This distinction between organic chemistry and inorganic chemistry was generally accepted until the year 1828, when the German chemist, F. Wohler, succeeded in preparing urea, an excretion product of animal organisms, by heating the inorganic salt, ammonium cyante, a substance which might be considered to be inorganic, or mineral.



This classic synthesis showed that the influence of a living organism was not necessary for the production of an organic substance—urea. As the science of chemistry was developed, it was soon found that a great many other so-called organic substances could be prepared in the laboratory by artificial methods and from materials of inorganic origin. Ultimately it came to be generally acknowledged that many of the processes which occur in animals and plants could very probably be carried out in the laboratory, and that the formation of an organic compound is probably not dependent at all on the help of any vital force. Today this difference between the two classes of compounds has been recognized as an imaginary one, and the terms "organic chemistry" and "inorganic chemistry," have, to a large degree, lost their original meanings. They do, however, serve to sub-divide the fields of chemistry into two groups which are characterized by their own special technique, and whose exploitation has led to products which have satisfied many human needs and produced a basis for important and basic chemical industries. The compounds of carbon are related to one another yet they differ widely in their general behavior from those of all other elements. Thus they form a very distinct group of compounds, and it is convenient to class them separately and to distinguish them by the term "organic." Organic chemistry, therefore, according to modern interpretation, is the chemistry of carbon compounds which can be created by application of the art of synthesis.

← **A View in a Modern Chemistry Laboratory.**
Photo Courtesy of Du Pont Company

Part Two

THE CHEMIST'S LABORATORY

THE chemist's work-room or laboratory has several special requirements if it is to be fully satisfactory. A room somewhat isolated to avoid interruption is desirable, especially if small children are around to stick inquisitive fingers into things. Good ventilation is necessary, and at least enough heat at all times to keep water solutions from freezing. While a well trained chemist seldom spills anything, and, in spite of popular opinion, almost never has an explosion, it is better to have the laboratory plainly and simply furnished so that an accidental splash will do no damage. A plain wooden floor is better than one covered with a carpet, and concrete or linoleum is still better. The work table may be of plain lumber, with the top waxed frequently to protect it. A sink and a supply of running water are quite essential, but if he lacks these the ingenious boy chemists will find a way to provide himself with running water from a pail fitted with a siphon and hose. And you never will get too many shelves, cabinets and drawers for storage.

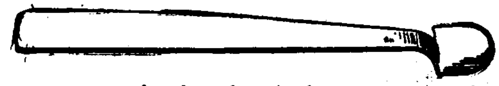
Now in picturing to you this ideal laboratory, we realize that few boys can have all this equipment at once. In fact, your Gilbert Chemistry Set has been designed to be as far as possible a complete laboratory in itself. But we feel sure you will enjoy it more if you can at least select for it a secluded corner in den or kitchen, or even in the woodshed, cellar, or attic where your apparatus may be left set up undisturbed and where there will be room to expand as you build or buy new equipment and supplies.

The Equipment and Its Use

Good technique can only be acquired by careful self-training. Learn how to use each piece of apparatus properly. Begin by having a place to keep each and every piece, and keep it clean and in its place. Be extremely careful not to contaminate your chemical supplies by getting even a trace of one chemical mixed in the bottle of another. And watch to keep your chemicals replaced as soon as the supply runs low.

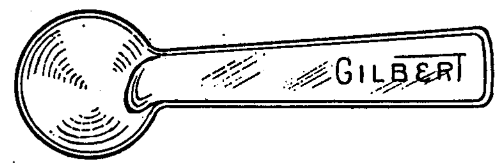
While you have not been furnished with dangerous and poisonous chemicals, they are not intended to be taken into the mouth. You should begin now to train yourself, never to taste anything in the laboratory, and to use caution in smelling any substance. Always exercise caution.

THE SPATULA AND MEASURE



One measure of a dry chemical means as much as can be held in the spoon-shaped measure (above). For transferring solid materials and for rough measuring the chemist uses a flat blade called a spatula. Your set has been equipped with an improved spatula having the flat blade at one end and a small spoon-shaped measure at the other. Even when made of corrosion-resistant metals, a spatula is soon corroded by chemicals unless you wash and dry it immediately after use. A roll of inexpensive paper toweling is invaluable for this and similar purposes in your laboratory.

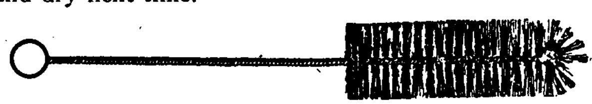
THE MEASURING SPOON



One teaspoonful of a dry chemical means as much as the spoon will hold tapping it lightly. The teaspoon is also used for heating solids.

TEST TUBE BRUSH

A test tube brush has been furnished to help clean test tubes. You will find that a little ordinary kitchen scouring powder on the brush will help greatly in cleaning your tubes. Always clean the test tubes immediately after you are through using them so they will be clean and dry next time.



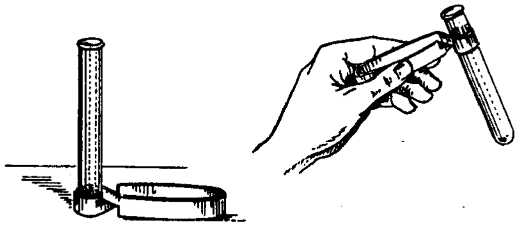
THE STIRRING ROD

A stirring rod (above) is a very convenient piece of apparatus for mixing a solution when dissolving a solid in a liquid. It is a solid glass rod, round at both ends. Always clean the rod with water and dry with a towel before using it in different solutions.

TEST TUBES

The test tubes in your set are not the miniature toy affairs sometimes put into chemistry sets, but practical test tubes made of especially strong, heat-resisting glass. Some skill is needed when heating liquids in a test tube to avoid sudden formation of steam which may throw some of the liquid out of the tube.

THE TEST TUBE HOLDER



It is recommended to always use the test tube holder when heating mixtures in a test tube because the tube will often become too hot for the fingers to hold. The Test Tube Holder can also be used as a stand.

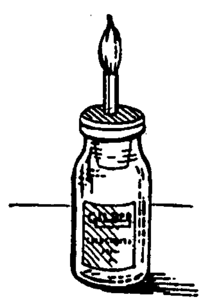
EXPERIMENT 1—Heating a Liquid in a Test Tube

Fill a test tube about one-third full of water and attach the test tube holder near the top of the tube. Hold the tube over the flame of the candle, keeping it in a slanting position as shown in the illustration above so that the heat strikes the side of the tube. Maintain a gentle rotary motion to promote smooth and steady boiling. Even with this precaution, do not point the open end of the tube toward yourself or any other person while you are heating it.

If a test tube has been heated empty or with dry solid materials inside, do not pour water or any other liquid into the tube until it has cooled.

Some pieces of equipment illustrated may not be contained in your set. However, they may be purchased separately if desired; see listing on page 80.

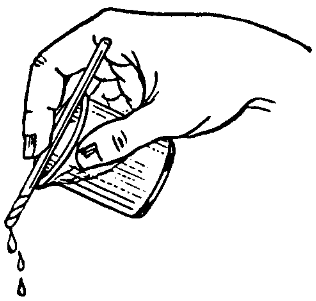
The alcohol lamp supplies enough heat for most laboratory purposes and provides a clean flame with less attention to adjustment than a gas burner.



BEAKER

The beaker is a straight sided glass container, generally made of Pyrex glass, used for mixing and heating quantities of material too large for a test tube. To avoid danger of breakage, a wire gauze or an asbestos mat should be placed between the free flame and the bottom of the beaker.

Although the beaker is provided with a lip for pouring, liquid sometimes runs down the outside of the beaker unless a glass rod is held across the lip of the beaker in the position shown in sketch.

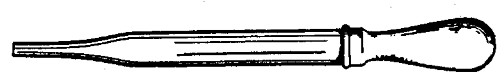


EXPERIMENT 2—To Pour Liquid from a Beaker

Fill a beaker nearly full of water and place a stirring rod across the top so that it rests on the lip of the beaker and the end extends a little beyond the lip. Now pick up the beaker with the rod held in position by the first finger as shown above and tip it slowly to pour. Notice that the liquid follows the glass rod and is much less likely to run down the side of the beaker than when no rod is used.

MEASURING LIQUIDS OR SOLUTIONS

The medicine dropper below is a very useful piece of apparatus for measuring out several drops of a liquid or solution. The dropper is filled by pressing on the bulb

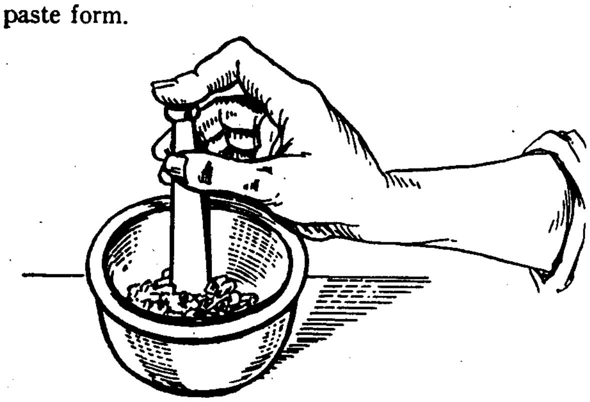


and inserting the open end into the solution. On relieving the pressure, the dropper fills with the solution.

Always wash out the dropper with pure water several times before using it with different solutions.

MORTAR AND PESTLE

The mortar and pestle are used for grinding lumps of solids to a powder or for mixing substances in dry or paste form.



Part III

CHEMICAL THEORY FOR BEGINNERS

KINDS OF MATTER

MATTER is found in different forms such as gas or vapor, liquid and solid, and many of the same substances of matter may be made to assume all the three different forms. Water for example, is of common occurrence in nature as a liquid, but if the temperature falls low enough, as in winter weather, water is chilled and finally becomes a solid which we recognize as ice. On the other hand if the temperature rises high enough, the water becomes invisible and turns into a gaseous state, which we utilize in the form of steam as a source of power in the steam engine. Water is an inorganic substance. The three forms of matter can also prevail among many organic compounds. For example, if one heats a piece of camphor gum in an evaporation dish or in a tin cup, he will observe the following phenomena: The solid particle of camphor turns to a clear liquid. In other words, the camphor, when first heated, passes from a solid into a liquid state by melting. On continued application of heat, the camphor will finally volatilize and become invisible. On leading the vapors of camphor over a cool surface, the organic vaporous material will solidify and deposit again as solid camphor. During all these physical changes there is no alteration in the chemical composition of the camphor. The same statement also applies to the inorganic material—water.

DIVISION OF MATTER

It is possible to divide inorganic matter into very minute particles. Glass, for example, may be broken up into fine fragments and even ground to a dust. In the manufacture of Portland cement, fine particles of inorganic matter are produced by intensive grinding of mineral substances. Every single particle of the dust of a pulverized substance represents the same composition as that of the original material before grinding. A wild animal running through a forest, like a fox or deer, may emit odorous particles in his travels that cannot be seen by any human eye, yet the hound can easily pick up the scent and follow them for miles hours after the fox or deer has passed a given point. The smell of a fragrant flower can be detected in a room even though just a trace of it may be present.

EXPERIMENT 3—Division of Matter

Equipment: Common salt, water and a small beaker or glass.

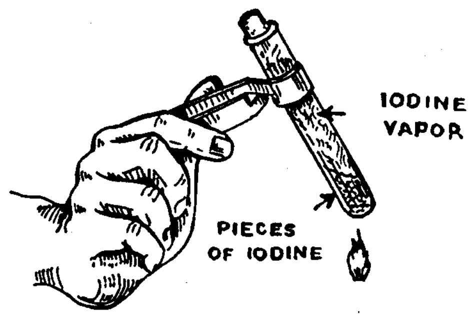
Operation: Take a pinch of common salt (sodium chloride) and dissolve it in water. Now taste the solution. You will be able to detect the characteristic taste of the salt in the solution, although you cannot see it with the human eye. Now add an equal volume of pure water to the salt solution. Mix well, and then taste the solution. You will find that you are still able to taste the salt. Dilute once more with an equal volume of water and then taste again the resulting solution.

Result: Now you can imagine the thousands of particles of salt there must be in your water solution to enable you to detect it by the taste and yet not to see it.

EXPERIMENT 4—Division of Matter

Equipment: Iodine crystals, test tube, cork, alcohol or candle flame.

Operation: Drop into a clean dry test tube a crystal



of iodine which may be obtained in a drug store, and insert a common cork lightly into the mouth of the test tube above. Hold the test tube over an alcohol or candle flame and warm gently. Watch the beautiful purple gas that is formed and observe how it creeps upward toward the mouth of the test tube.

Result: Here you have another demonstration of extreme divisibility; the solid particles of iodine having been volatilized by heat and divided into thousands of small gaseous particles of iodine which are visible to the eye.

INORGANIC AND ORGANIC MOLECULES

Molecules are small particles of matter. The three foregoing experiments have led chemists to the conclusion that ordinary matter is composed of particles and that these particles are so small and so minute in size that they cannot be detected with the human eye or even under the magnification of the most powerful microscope. As we proceed with our experiments, we will speak of these different particles of matter as molecules, and an important law of chemistry may be stated as follows: The smallest particle into which matter can be sub-divided without changing its chemical nature, is a molecule. Having adopted a name—"molecule"—for the smallest particle of any substance which has all of the properties of a whole substance, it is now easier to explain the physical difference between solids, liquids and gases.

MOBILITY OF MOLECULES IN SOLIDS, LIQUIDS AND GASES

In a solid substance, the chemist conceives molecules as being firmly held together. They are not free to move about to any great extent, and have fixed positions. For this reason a solid substance retains its form or shape. Many substances assume definite crystalline form of great beauty. In crystals we have evidence to support the conclusion that the molecules are arranged in an orderly fashion and that it is due to this orderly arrangement of the molecules of both inorganic and organic substances that crystals assume such beautiful symmetry. We have no better illustration of variations in crystalline form than that revealed by snow flakes. In other words, the face of a crystal is composed theoretically of a layer of orderly arranged molecules. In a liquid, on the other hand, the molecules are free to move about among themselves much more freely than in a solid. For this reason liquids flow and assume the shape of the receptacle in which they are inclosed. Solids can be broken up and ground into smaller parts, while a liquid will withstand great pressure and cannot be divided by frictional force. Gas molecules are assumed to be entirely separate and mingle with each other and fly apart very widely when they are allowed to leave the space in which they are confined. In other words, the molecules are mobile, and due to this property, we speak of a gas as being volatile and easily diffused. The molecules of a gas move about much more rapidly than those of a liquid or a solid. To illustrate, the walls of a football are kept pushed out by the constant pressure and hammering of millions of air molecules that are enclosed within the walls of the football. Also the pressure within a steam boiler is caused by the molecules of water in the form of water vapor or steam which press constantly against the sides of the boiler.

PHYSICAL PROPERTIES OF MATTER, MASS AND VOLUME

All matter occupies space. The amount of space occupied is called volume. A bushel, a quart, a liter, each describe a measure of volume of space occupied by a particular substance. The space occupied by a known weight of solid liquid, or gas varies according to the density of the material. By specific gravity we mean the relation between the volume and the weight of a substance. Water is used as the standard, that is, 1 cubic centimeter of water at 4 degrees Centigrade weighs 1 gram. Thus, 1 cubic centimeter of a substance which weighs 5 grams has a specific gravity of 5.

The specific gravity of a substance often enables us to tell whether or not that substance has been adulterated. For example, a certain oil may have a specific gravity of 2. If the specific gravity when measured is 1.7, we know the oil has been adulterated with some other substance and is probably inferior in quality.

EXPERIMENT 5—To Demonstrate the Specific Gravity of Liquids

Equipment: Fresh egg, glass of water, spoon and common salt.

Operation: Place the fresh egg in a glass full of water and notice that it sinks to the bottom. Now remove the egg and dissolve two spoonfuls of salt in the water. Then put the egg into the salt solution and notice that it floats.

Result: The reason why the egg floated in the salt solution was because we increased the specific gravity of the water beyond that of the egg. The specific gravity of a solution is always greater than that of the pure liquid. This is the reason why it is easier to swim in salt water than in fresh water.

MALLEABILITY

By malleability we mean that a metal can be rolled out or flattened easily. This property is characteristic of many metals, particularly platinum and gold. Gold can be rolled into leaves so thin that it would take 300,000 of these leaves to make the thickness of one inch. Iron is a very malleable metal when hot, and can be moulded into many shapes and forms available for human needs. Many thousands of men are employed daily in large steel mills in this country, rolling steel for railroads, girders for buildings and bridges, sheet metal for roofs, steel plates for oil tanks, and many other applications made possible by man's ingenuity.

DUCTILITY

Ductility means that a metal can be drawn out into a fine thread or wire. Iron, copper, platinum, gold, and many other metals are characterized by this property. Ductile metals are especially important to the electrical industry. It is interesting to know that the strength of some metals is increased by drawing them out into small wire and for this reason a drawn wire is stronger than an ordinary piece of metal of the same dimensions. Large cables made by twisting iron wire together are much stronger than solid iron rods of the same size. Cables of this type are used in the construction of suspension bridges, of which the George Washington Bridge, extending over the Hudson River, near New York City, is an excellent example.

BRITTLENESS

When materials can be easily crumbled and broken, we call them brittle. Excellent examples of a few substances which are brittle are ordinary glass, chalk, ice, egg shells, and almost all rock.

ELASTICITY

When a material can be stretched and it will return to its original form, we call it elastic. The molecules of the substance apparently tend to return to the places they formerly occupied before the change. Substances possessing this characteristic property are spoken of as being elastic substances. The outstanding illustration is ordinary rubber. In stretching a piece of rubber we disturb the arrangement of the molecules of this material.

HARDNESS

A solid is said to be harder than another solid when it will scratch or make a mark on the other substance. For example, a diamond will scratch, and even cut glass, because it is harder than glass. You can cut ordinary metallic lead with a knife blade because it is softer than the metal of the knife.

FUNDAMENTAL PHYSICAL CONSTANTS WIDELY USED BY THE CHEMIST

Freezing Point or Melting Point

The melting point of a solid is that temperature at which it is converted to a liquid. It is a physical constant which is characteristic of all substances that are chemically pure. For example, ice has a melting point of 0° C., or 32°F. above which temperature it is slowly converted into water. The freezing point of water (liquid) corresponds to the melting point of ice (solid).

Boiling Point

This is the temperature at which a liquid is converted into a vapor. The constant is dependent upon the atmospheric pressure existing at the time of distillation of the liquid. For example, water will boil at a lower temperature at a high altitude, as on a mountain top, where the pressure is less than it will at sea level. We have many liquids that cannot be distilled at ordinary pressure without decomposition, but if such substances are distilled in a vacuum, they can be easily purified without loss. The instrument which is used by the chemist for determining melting points and boiling points is a thermometer, and two thermometer scales find application in chemical practice—the Fahrenheit and the Centigrade scales. According to the Fahrenheit scale, water freezes at 32° above the zero point, and boils at 212°. According to the Centigrade scale, the freezing point of water is 0°, and the boiling point is 100°. In chemical practice and also for recording scientific data in chemical publications we today make use of constants as recorded by both the scales.

THE ELEMENTS AND CHEMICAL CHANGE

All of the changes which take place in chemical reactions are based on the actions of certain substances which the chemist calls "elements." A boy or girl should, therefore, first understand the correct meaning of this term if he or she is to do the experiments intelligently. By elements the chemist means those substances which he is not able to break up into simpler substances. An element may be a solid, such as copper or iron; a liquid, such as mercury; or a gas, such as oxygen or hydrogen. Some elements can exist in all of the three different physical forms. For example, ordinary iron is a solid, but will become a liquid if heated hot enough, or if heated at a very high temperature, it will become a gas. Mercury can be heated to form a gas, or it can be cooled to form a solid. The gas, oxygen, can be cooled to a liquid, or when cooled still further, to a solid. In all these physical changes of iron, mercury and oxygen, brought about by changes of temperatures, we have not altered the elementary nature of these three substances.

More than ninety of these elements have been discovered to date, and we know that they combine in millions of different ways to form every substance that we know on our earth. When elements combine with each other they form what the scientist calls "compounds." When elements are converted into compounds, they lose their chemical identity, and we create new substances possessing different properties. Such phenomena we explain as chemical changes, which the chemist also calls chemical reactions. For example, sulphur is an element and oxygen is an element. When the sulphur is heated it attracts the element oxygen, and undergoes a chemical change. In other words, sulphur and oxygen undergo a chemical reaction to form a compound which is a gas-sulphur dioxide. This gas is used commercially as a bleaching agent, and also as a cooling agent in electric refrigerators. Therefore we may say that chemistry is the science dealing with the study of the elements and compounds, their properties, their chemical changes, and reactions.

THE ATOMIC THEORY

Under the heading—Inorganic and Organic Molecules—we stated that a molecule is the smallest particle into which matter can be divided without changing its chemical nature. This definition of a molecule applies to both elements and compounds, and just as long as one does not subject an element or a compound to an experimental condition which leads to a chemical change or reaction, the substance is said to retain its original form.

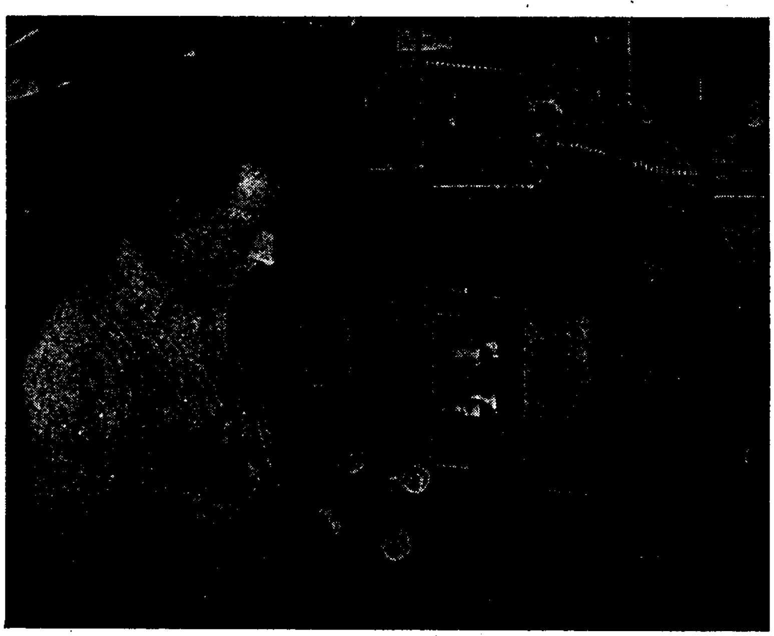
When a chemical change or reaction takes place, the molecules of an element or compound are destroyed and new molecules are formed. In order to satisfactorily describe this situation the chemist developed a fundamental theory of chemistry known as the Atomic Theory. The chemist's explanation of chemical change or chemical reaction on the basis of the Atomic Theory is as follows: He assumes that there are particles in matter that are much smaller than the molecules to which we have previously referred. This division of molecules takes place under conditions favoring a chemical change, and when molecules break up into these new imaginary particles the substance becomes very unstable and very reactive. The chemist has named these invisible and imaginary units "atoms" and when compounds are formed by interaction of elements or otherwise, it is the atoms that are involved, and new compounds are formed. In other words, molecules of elements or compounds lose their physical identity when they react chemically to form new substances, and are converted first into atoms. These atoms are very unstable and cannot exist alone, and consequently seek other atoms to make new molecules. The result, in chemical sense, is a chemical reaction, with formation of a new substance having entirely different chemical and physical properties than those of the original substances brought together.

Therefore, as a result of the above reasoning, chemists have adopted the law of the atom to guide them in their chemical reasoning. It is as follows: *An atom is the smallest particle into which matter is divided in chemical changes.* Chemistry then has a much broader meaning when we work according to such a theory, and may be defined as the science of atoms, and how atoms combine with each other to form compounds.

The interaction of sulphur and oxygen may now be explained as follows: When these two elements combine with each other it is not the molecules of sulphur and oxygen that react, but their atoms. The atoms on being formed cannot exist alone, and attract each other, leading to the formation of a new compound in which the properties of sulphur and oxygen have been lost entirely. An atom of sulphur (S) combines with two atoms of oxygen (O_2) to form a new substance, or gas (SO_2). The combination of two different atoms in this case leads to the formation of a new molecule.

The Modern Conception of the Atom

Great advances in the science of chemistry and physics have been made by scientists in the last thirty-five years. New theories regarding structure have been advanced which have led to a complete revolution of previous concepts. According to the new theories of the



THERMOPLASTIC MOLDING MACHINE AND PRODUCT.

All plastics are chemical derivatives of such basic materials as coal, cellulose and petroleum. After a plastic material, such as Bakelite, has been manufactured by chemical process, it can be molded under conditions of heat and pressure into any one of numerous shapes and sizes depending upon the mold used. The machine illustrated is used at the Gilbert plant to mold toys and household appliance parts.

structure of matter, the atom is not the smallest particle which functions in a chemical change, but we now conceive the existence of a still smaller unit than the atoms —namely, the *electron.* This new conception of chemical change is electrical in nature and the electron is the unit of electrical charge. These newer ideas are constantly undergoing modification and are far too advanced and theoretical to be comprehended by boys and girls for whom this manual is written, but it is not out of place to record here some of the conclusions that have been accepted and which are influencing the development of modern chemical reasoning. The modern concepts regarding the structure of atoms may be stated briefly as follows:

1. Each of the 92 elements which we believe to constitute all matter is made up of atoms.

2. Atoms in turn consist of units which we call protons and electrons.

3. Protons are units charged positively and electrons are units charged negatively.

4. Every atom contains an equal number of positive and negative charged units.

5. The atom owes the greater part of its weight to the positively charged protons.

6. All the protons and a part of the electrons of each atom are concentrated within a relatively small space at the center of the atom. The remaining electrons are located in layers or "shells" outside of the nucleus. The nucleus may be conceived therefore as the sun of a planetary system around which rotate the planets (electrons).

LIST OF ELEMENTS

	Symbol		*Symbol*		*Symbol*
Aluminum	Al	Holmium	Ho	Rhenium	Re
Antimony	Sb	Hydrogen	H	Rhodium	Rh
Argon	A	Indium	In	Rubidium	Rb
Arsenic	As	Iodine	I	Ruthenium	Ru
Barium	Ba	Iridium	Ir	Samarium	Sm
Beryllium	Be	Iron	Fe	Scandium	Sc
Bismuth	Bi	Krypton	Kr	Selenium	Se
Boron	B	Lanthanum	La	Silicon	Si
Bromine	Br	Lead	Pb	Silver	Ag
Cadmium	Cd	Lithium	Li	Sodium	Na
Calcium	Ca	Lutecium	Lu	Strontium	Sr
Carbon	C	Magnesium	Mg	Sulphur	S
Cerium	Ce	Manganese	Mn	Tantalum	Ta
Cesium	Cs	Mercury	Hg	Tellurium	Te
Chlorine	Cl	Molybdenum	Mo	Terbium	Tb
Chromium	Cr	Neodymium	Nd	Thallium	Tl
Cobalt	Co	Neon	Ne	Thorium	Th
Columbium	Cb	Nickel	Ni	Thulium	Tm
Copper	Cu	Nitrogen	N	Tin	Sn
Dysprosium	Dy	Osmium	Os	Titanium	Ti
Erbium	Er	Oxygen	O	Tungsten	W
Europium	Eu	Palladium	Pd	Uranium	U
Fluorine	F	Phosphorus	P	Vanadium	V
Gadolinium	Gd	Platinum	Pt	Xenon	Xe
Gallium	Ga	Potassium	K	Ytterbium	Yb
Germanium	Ge	Praseodymium	Pr	Yttrium	Y
Gold	Au	Protactinium	Pa	Zinc	Zn
Hafnium	Hf	Radium	Ra	Zirconium	Zr
Helium	He	Radon	Rn		

LARGE CHEMICAL INDUSTRIAL PLANT

Many of the experiments perfected on apparatus in a chemistry laboratory are later put into large scale production using an elaborate installation such as illustrated above.

Part IV

CHEMICAL CHANGES

IN general, there are four kinds of chemical changes which we need to consider in our elementary study, and every chemical reaction described in this book will come under one of these changes. We will illustrate by experiments the four important types. First, we have that of Direct Union, or the combining of two elements to form a compound; second, we have Decomposition or Degradation, which means the breaking down of a compound into its elements or into simpler substances; third, Double Decomposition or the exchange of elements in two or more substances to form new compounds, and fourth, Substitution or Replacement, a reaction in which one element takes the place of another in a compound, the substituted element being set free.

CHEMICAL CHANGE THROUGH DEGRADATION BY HEAT

EXPERIMENT 6—Decomposition of Sugar

Equipment: Sugar, measure, test tube, test tube holder and alcohol lamp or candle.

Operation: Put two measures of granulated sugar in a dry test tube and cautiously heat the tube over an alcohol lamp. Note the change in color of the sugar and its tendency to liquify. Watch the molten sugar and continue heating until it begins to char and turn black. Here we have a simple demonstration of degradation of an organic substance by heat. A chemical change has taken place by heating, leading to destruction of the sugar molecules with formation of water molecules and ordinary carbon. The identity of the carbon is disguised in the original, colorless sugar molecule, but is revealed when the sugar molecule undergoes decomposition.

Result: It is a very common property of many organic substances to decompose on heating. All organic animal matter decomposes on intense heating to yield free carbon. This is known as *carbonization.*

CHEMICAL CHANGE BY DOUBLE DECOMPOSITION

EXPERIMENT 7—Action of Ferric Ammonium Sulphate on Calcium Oxide

Equipment: Ferric ammonium sulphate, calcium oxide, water, test tube, and measure.

Operation: In a test tube of cold water add ½ measure of ferric (iron) ammonium sulphate. Place the thumb over the mouth of the test tube and shake to dissolve the solid. Now add ½ measure of powdered calcium oxide and shake again. A reddish brown precipitate is formed.

Result: The iron of the ferric ammonium sulphate changes places with the calcium of the calcium oxide to form calcium sulphate and ferric (iron) hydroxide, which is insoluble in water, and appears as a reddish-brown precipitate.

THE ATMOSPHERE

The air enveloping our earth is essentially a mixture of two elements—oxygen, which comprises about one-fifth of the air by volume, and nitrogen, four-fifths by volume. Actually, oxygen and nitrogen comprises 99 per cent of the atmosphere by volume at sea level, the remaining one per cent being made up of a mixture of rare gases, namely argon, helium and neon, mixed with traces of hydrogen and carbon dioxide. Oxygen is considered to be the most abundant of all the elements and most widely distributed. Eight-ninths of water by weight is combined oxygen. Such common materials as sandstone, quartz, limestone or marble, common brick, granite, clay and cement contain fully one-half their weight of oxygen. About two-thirds of the weight of the human body is oxygen. The total weight of oxygen in the land, the water, the atmosphere, and in living organisms may be regarded as very nearly equal to the combined weights of all the other elements.

OXYGEN

Oxygen of the air plays a very important role in our everyday life. We breathe this element into our lungs and from these organs it is carried by the blood-stream throughout the body to combine with the waste products of body metabolism. The oxygen carried in the blood stream oxidizes the waste materials forming carbon dioxide. This carbon dioxide is carried back to the lungs by the blood and exhaled into the air in breathing. It is very essential, therefore, that we breathe in fresh air if we wish to enjoy good health. Pure oxygen is a very active element, and if it were not for the fact that the oxygen of the air is diluted with nitrogen gas, a very inactive substance, the world would soon burn up and all living organisms be destroyed.

Oxygen is said to support combustion, but the gas will not burn itself. A fire will not burn unless air rich in oxygen is constantly supplied so that the fuel can have plenty of oxygen to support combustion. Substances cannot burn without oxygen.

THE OXYGEN MASK AND CLOTHING USED IN STRATOSPHERE FLYING

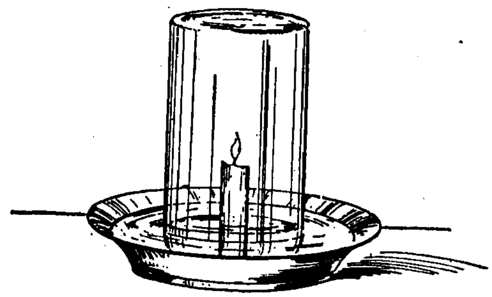
The picture below shows what the United States Army flier wears when flying at 40,000 feet. At this altitude in a "Flying Fortress" he wears regulation underclothes, trousers and blouse; an outer suit of leather and under it a processed "chicken feather" suit for extra warmth; helmet and goggles, two pairs of gloves—one of silk next to the hands—and another of leather. The rubber tube connects to an oxygen tank.

OXYGEN EXPERIMENTS

EXPERIMENT 8—To Remove Oxygen from Air

Equipment: Candle, water, wash pan, fruit jar.

Operation: Place a candle in the center of an ordinary wash pan by allowing a little of the melted wax of the candle to fall on the pan to stick the candle firmly. Then pour into the pan two inches of water (Fig. 18). Light the candle and place a fruit jar over it. Be sure that the jar is high enough so that the flame does not come too close to the top of the jar.



Result: You will notice that very soon the flame grows dim and finally dies out, the oxygen of the air within the jar having been entirely used up. Notice that the water begins to rise inside the jar and stands at a higher level than the water in the pan. This is because the oxygen was removed, forming a partial vacuum which drew the water up into the jar. The oxygen in the jar united with the carbon in the flame to form carbon dioxide, a gas which dissolves in water, and with hydrogen to form steam, which condenses to water. The gas remaining in the jar is essentially nitrogen.

HYDROGEN PEROXIDE

Hydrogen peroxide is a chemical with which every boy and girl should be familiar. It is found in every "First-Aid Cabinet" and is a good antiseptic. Its antiseptic properties are based on an extra oxygen atom which it gives up very easily. It is an unstable compound and frequently loses its strength due to the lowering of the oxygen content as a result of being kept on the druggist's shelf. Sometimes, this chemical is sold as a high grade peroxide solution when it actually never contained the required amount of excess oxygen. *Hydrogen peroxide solution should never be taken by mouth.*

EXPERIMENT 9—Behavior of Hydrogen Peroxide Toward Blood

Equipment: Blood, watch glass, hydrogen peroxide, glass rod.

Operation: Place a drop of blood on a small watch glass and then add a small quantity of hydrogen peroxide solution. Mix the blood with the peroxide solution by stirring with a glass rod. Notice that the peroxide is decomposed with evolution of bubbles of oxygen. At the same time the red color of the blood is lost.

Result: The blood is destroyed by the action of the generated oxygen. Many forms of bacteria are killed by contact with hydrogen peroxide.

COMMERCIAL USES FOR HYDROGEN PEROXIDE

Bleaching: Cotton, linen, rayon, silk and mixed fibres are bleached with hydrogen peroxide. Other materials bleached with this chemical are; straw, fur, pearl bone and plastic buttons, wood, vegetable oils, beeswax and sponges.

Hydrogen peroxide is used as an oxidizing agent in the manufacture of dyes; sulphur chemicals for the rubber industry and for manufacturing medicinal preparations.

OXIDATION

Some compounds contain much oxygen and under suitable conditions readily give up part or all of their oxygen to other compounds. Such substances are called by chemists—*oxidizing agents.* In other words, by oxidation is meant the union of a substance with oxygen. During intense chemical oxidation heat is evolved and sometimes light.

Oxidation is an important chemical process. We obtain heat to warm our homes in winter, and power to run machinery in our factories by burning (oxidizing) wood, oil and coal. Our houses are lighted by electricity that is generated by machinery run by burning (oxidizing) fuel. When gasoline is ignited and exploded in the cylinders of an automobile engine, the gasoline suddenly unites with the oxygen of the air which has been drawn into the cylinders. An ordinary furnace in the home is an oxidizing machine, and even man and all living organisms (animals) are likewise active oxdizing machines. By means of heat obtained by the process of oxidation metals can be melted. By power obtained from the process of oxidation our buildings are cooled and water can be frozen to artificial ice. Ice machines are in common commercial use today, and play an important part in the preservation of health. In fact, human life and all human activities depend upon some form of chemical oxidation and without oxygen our lives and those of the animals and plants would cease.

REFRIGERATION AND AIR CONDITIONING

Reversed refrigeration, or the process of heating a house in winter by means of the same equipment for cooling it in summer, is an idea which is being carefully investigated today by engineering concerns interested in the manufacture of air-conditioning apparatus. In summer, air-conditioning equipment absorbs heat from inside the house, takes it to the outside and discards it, just as an electric refrigerator takes up heat from inside the food compartment and releases it into the room. In winter the cycle is reversed. Heat is taken up from out doors and brought inside to warm the house. Coils inside the house, which formerly absorbed heat, become radiators, while the outside coils, instead of throwing off heat, absorb warmth from the outside air. There are many problems yet to be solved before reversed refrigeration equipment is commercially successful, but the principle is constantly finding new applications and a promising future is ahead for new and important commercial developments.

EXPERIMENT 10—Fire Ink

Equipment: Potassium nitrate, test tube, candle, small brush.

Operation: Place ½ spoonful of potassium nitrate in a test tube and add ½ inch of water. Warm over the candle for a minute to dissolve all the material.

Now write with this liquid upon some unglazed or porous paper, using a clean pen or a small brush. Be sure that the strokes are heavy and all lines are connecting. After the lines are thoroughly dry apply a lighted match or better a glowing spark to some of the writing. Blow out any flame that may result.

Result: If properly done, the spark will travel along the lines where the liquid has been applied, leaving the rest of the paper untouched. The potassium nitrate is a strong oxidizing agent.

This experiment is very mystifying when performed in the dark. The best results are obtained by using soft paper, and by making the lines heavy and connecting.

EXPERIMENT 11—How to Make a Fuse

Equipment: Potassium nitrate, test tube, spoon.

A very good fuse can be made by soaking a cotton string in a solution of potassium nitrate or saltpetre for a few minutes and then allowing the string to dry. Allow the string to be suspended while drying. You can time your fuse by using the proper length of string. The nitrate solution is prepared by dissolving ½ spoonful of potassium nitrate in a test tube containing ½ inch of water, and then shaking until well dissolved.

EXPERIMENT 12—Oxidation of an Element by Means of a Nitrate

Equipment: Potassium nitrate, sulphur, spoon, measure, alcohol lamp or candle.

Operation: Heat on your spoon one measure of potassium nitrate until the salt is molten. Then drop a pinch of sulphur into the molten potassium nitrate and notice the sudden flash.

Result: The sulphur will be oxidized by the oxygen from the potassium nitrate to form sulphur dioxide. Note the odor of the burning sulphur. In the preceding experiment the burning of cotton fuse leads to the formation of carbon dioxide, an odorless gas, while sulphur burns under similar conditions to give sulphur dioxide, having a characteristic penetrating odor.

EXPERIMENT 13—Oxidation of Spices

Equipment: Spices, potassium nitrate, measure, alcohol lamp or candle, spoon.

Operation: Mix together 4 parts cinnamon, 3 parts of allspice, and 5 parts of ground cloves and grind together. Now add 8 parts of potassium nitrate to the above mixture but do not grind. Place some on a spoon and warm.

Result: Notice the odor—much like perfume. Ignite this mixture and a beautiful shower of colored fire will be the result.

EXPERIMENT 14—Suffocating a Burning Candle

Operation: Attach a short piece of a candle to a cork and float this on water in a tin pan about 2 or 3 inches deep. If the candle is top heavy fasten a nail or small iron weight to the underside of the cork. Light the candle and cover it with an inverted fruit jar placed with its mouth on the base of the pan and under the water surface. As the candle burns it becomes paler and finally the flame dies out.

Result: You will notice that the water level in the jar is higher than at the beginning of the experiment. Some of the oxygen of the air has disappeared to support the combustion of the candle. Close the mouth of the fruit jar with a sheet of cardboard and set upright on the table.

EXPERIMENT 15—Testing for Carbonic Acid Gas

Insert a lighted taper into the jar from the preceding experiment. The flame will be extinguished. The carbon dioxide in the jar will not support combustion.

EXPERIMENT 16—Burning Sulphur

Operation: Repeat the candle experiment (14) using some sulphur. Place the sulphur on a small tin lid resting on a cork. Ignite the sulphur and burn under the fruit jar. After the flame is extinguished set the fruit jar upright on a table and notice the color of the gas. Also suspend a moistened blue litmus paper inside the jar.

Result: Sulphur burns to form sulphur dioxide. Sulphur dioxide is soluble in water, forming an acid. This will turn the blue litmus paper red.

EXPERIMENT 17—Burning Iron

Equipment: Iron, cellophane, test tube, water, pan.

Operation: Collect a few clean pieces of iron free from all rust. Tacks and small nuts and bolts are suitable. Wrap these in a piece of sheet cellophane and force firmly into the bottom of a medium sized tube. Then invert this tube in a pan of water, support it by a clamp and let stand for several hours.

Result: The water will gradually rise to a higher level in the test tube, showing that air has been used up. The iron is slowly oxidized in contact with moist air and is changed by the oxygen of the air forming iron oxide. Examine the fragments of iron and note their appearance.

EXPERIMENT 18—Protecting Iron from Oxidation

Operation: Repeat the preceding experiment, but use fragments of iron which have been coated with collodion or some other material impervious to moist air.

Result: The water will not rise in the jar. You will observe that there is very little tendency here for the iron to undergo oxidation. Paint serves to preserve iron from oxidation or corrosion.

EXPERIMENT 19—Oxidation in the Body

The changes taking place in our body are similar to the preceding changes of burning candle and sulphur. Our bodies consist of complex carbon compounds, some of which actually contain sulphur. When we breathe we inhale air through the lungs and here the oxygen of the air is picked up by the blood and carried where needed in the body. A burning process actually takes place internally, and the products of combustion are expelled through the lungs. Blow through a glass tube into a test tube of lime water. What happens? A white precipitate of calcium carbonate is formed, showing the presence of carbon dioxide.

EXPERIMENT 20—Combustion of Charcoal *

Equipment: Potassium nitrate, charcoal, test tube, alcohol lamp or candle, measure.

Operation: Heat three measures of potassium nitrate in a dry test tube over your alcohol lamp until the salt crystals liquify. Holding tube vertically, drop some specks of charcoal into the molten salt.

Result: The charcoal will take fire immediately. The nitrate furnished oxygen to burn carbon.

EXPERIMENT 21—Burning Sulphur

Repeat the preceding experiment using sulphur instead of charcoal. This will burn in the presence of potassium nitrate, forming sulphur dioxide, which will be detected by its characteristic odor.

EXPERIMENT 22—Burning Paper

Repeat the above experiment with fragments of dry filter paper fiber.

EXPERIMENT 23—Oxidizing Copper

Insert a piece of copper wire into some molten potassium nitrate. Note the change after continued heating over your alcohol lamp or candle.

EXPERIMENT 24—Oxidizing Iron

Repeat the above experiment using a piece of iron wire.

EXPERIMENT 25—Oxidizing Aluminum

Repeat the above experiment using a small piece of aluminum foil or wire.

EXPERIMENT 26—Oxidizing Silver

Repeat the above experiment using a piece of polished silver metal.

EXPERIMENT 27—Asbestos Insulation

Repeat the experiment using asbestos fibers. What happens?

EXPERIMENT 28—Cotton Fiber

Repeat the above experiment with cotton fiber. Which is the best fire insulating material, asbestos or cotton fiber?

PROTECTION AGAINST FIRE

Intensive oxidation of combustible material leads to generation of heat and as a final result, to a conflagration. Cloth, wood, paper and other substances may be rendered fireproof by treating them with the proper chemicals. Ammonium chloride or sal ammoniac (NH_4 Cl) is a cheap salt which can be used for this purpose. The article to be fireproofed is dipped or soaked in a strong aqueous solution of ammonium chloride and then dried. When such treated material is heated, the ammonium chloride is decomposed with liberation of ammonia and hydrochloric acid, and a fire is prevented, as neither the ammonia nor hydrochloric acid will support combustion. When they are generated they smother the flame and act as a fire extinguisher. The curtains and scenery of theatres, and tapestries in public buildings are fireproofed as a protection against fires. Fireproofed wood finds important commercial use as a construction material. Such chemically treated materials cannot be set on fire by sparks or flames. Tin salts are frequently used in fireproofing.

EXPERIMENT 29—To Fireproof Cloth or Paper

Equipment: Ammonium chloride, test tube, paper or linen, spoon, water.

Operation: Take a strip of paper or linen and immerse it in a solution made by dissolving one teaspoonful of ammonium chloride in a test tube 1/3 full of water. When the paper or linen is dry, try to light it with a match.

Result: You will see that it burns while held in the flame but will go out just as soon as the flame is removed.

EXPERIMENT 30—To Fireproof Wood

Wood also is treated sometimes with a strong ammonium chloride solution. Another way to fireproof wood is to paint it with water glass solution.

Equipment: Sodium silicate (water glass), match.

Operation: Holding a match by the head, dip the other end in a sodium silicate solution (water glass). Allow the coating to dry for twenty minutes, then light the match.

Result: The flame will go out just as soon as it reaches the portion that has been dipped in the water glass solution.

EXPERIMENT 31—Fire-Proofing Cloth

Cloth or other inflammable substances may be fireproofed by treating them with chemicals which when heated give off vapors that smother the flame.

Equipment: Ammonium chloride, test tube, measure, cotton cloth, glass.

Operation: Dissolve 12 measures of ammonium chloride in a test tube 1/3 full of water. Put a piece of cotton cloth 2 or 3 inches square in the bottom of a glass and pour the liquid in the test tube over it. Stir the cloth around until it is wet through, then let it dry and try to light it with a match.

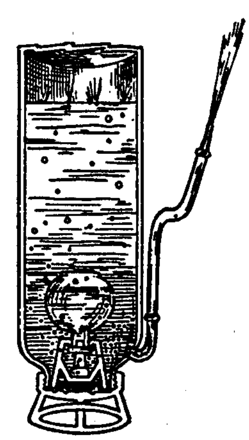
Result: You will find that it will burn while held in the flame, but just as soon as the match flame is removed it will go out.

FIRE EXTINGUISHERS

Since carbon dioxide neither burns nor supports combustion, it is used in fire extinguishers. You are all familiar with the hand fire extinguisher. Remember the instructions: "To operate, turn upside down." This apparatus is really no more nor less than a huge siphon, similar to the kind used for carbonated or seltzer water.

In the large vessel is a dilute solution of sodium carbonate. In the bottle is strong sulphuric acid. The cork is made of lead and fixed so that when the tank is turned upside down it falls from the bottle just enough to allow the sulphuric acid to trickle through slowly (see below). (A) shows an interior view of the fire extinguisher: (B) shows what happens when the extinguisher is turned upside down.

A



B

The acid liberates carbon dioxide from the sodium carbonate (just as you liberated carbon dioxide from sodium carbonate with tartaric acid.) The free gas creates a high pressure in the tank, causing large quantities of gas to dissolve in the water and forcing out a stream of gas and water. Carbon dioxide is heavy and surrounds the flame with a blanket of unburnable gas, which prevents access of oxygen thus smothering the flame.

Quite recently a new style fire extinguisher has come into use. The container is built like a hand pump. In the pumplike arrangement is a liquid called carbon tetrachloride. This is a cheap and rather marvelous liquid. It volatilizes, that is, it is converted into a gas as quickly as alcohol or gasoline but it does not burn. You are familiar with the explosive power of gasoline. Here is a vapor which defies all attempts to burn it. It is a non-combustible substance.

When directed at the base of the flame the spray of carbon tetrachloride quickly forms a gas which surrounds the burning material like a blanket and prevents the access of oxygen. Without oxygen there can be no combustion, so the fire is smothered.

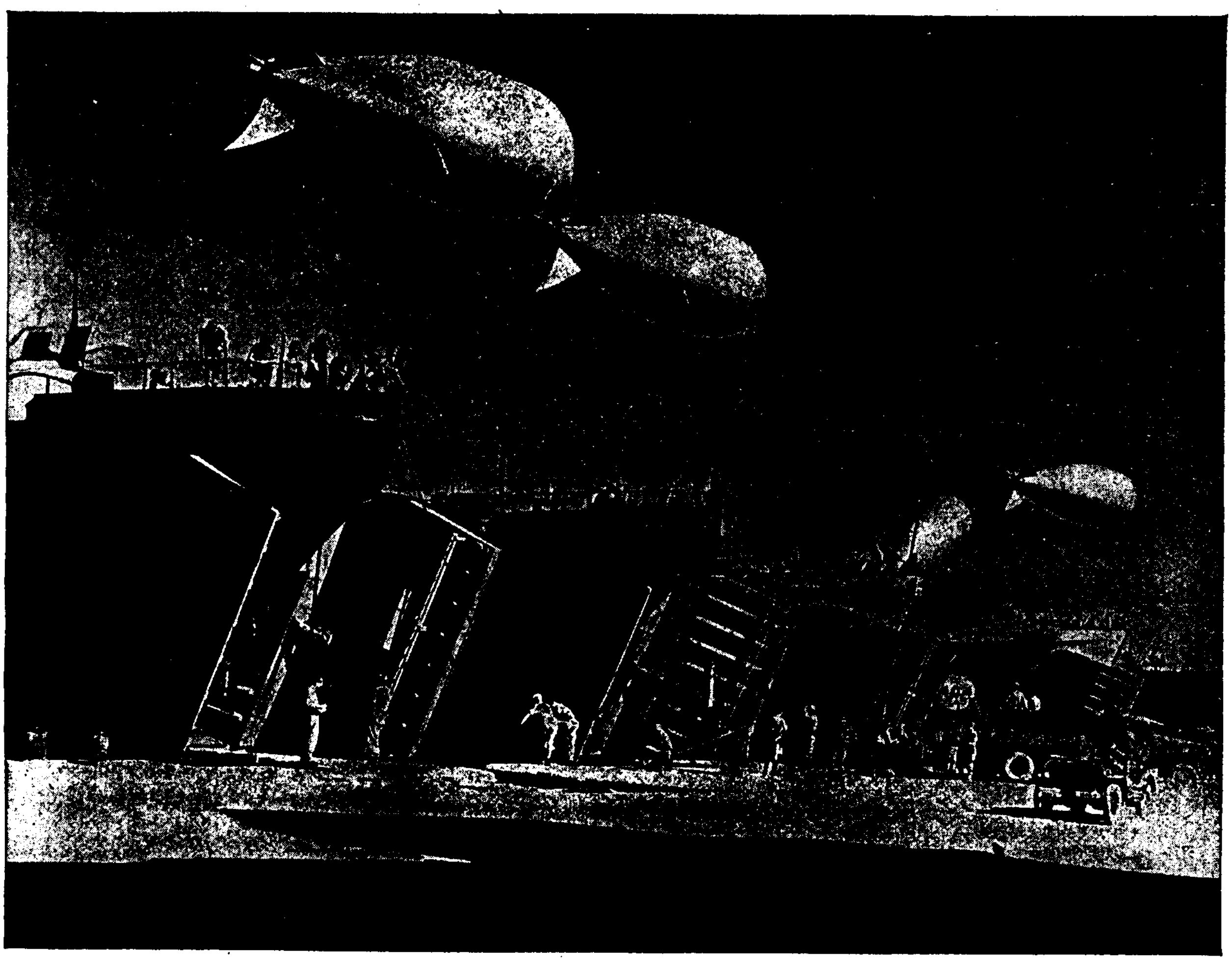
Carbon tetrachloride is also a splendid solvent. It is used to remove grease and paint spots from clothing and to clean white leather. You have probably noticed advertisements of spot removers which do not burn. All of these contain carbon tetrachloride. Carbon tetrachloride is a safe chemical to have around the house.

EXPERIMENT 32—A Hand Grenade Fire Extinguisher

Equipment: Crude calcium chloride, salt, water, test tubes, bottles, spoon.

Operation: Take 2 test tubes of crude calcium chloride, 2 spoonfuls of common salt and 1 cup of water and place in thin bottles.

Result: In case of fire, a bottle of this mixture thrown so that it will break near the flames will put the fire out. This mixture is better and cheaper than most of the grenades sold for the purpose of fire protection.



LST boats loaded with troops and equipment for invasion of Italy, at docks of Palermo, Sicily. Barrage balloons filled with Helium Gas form protective covering against enemy aircraft.

Part V

INORGANIC CHEMISTRY AND THE INDUSTRIES

HYDROGEN

HYDROGEN is one of the most interesting and important gases with which the chemist has to deal. It is, however, a dangerous gas to be used for experimentation by unskilled workers, and, therefore, the author of this manual does not recommend its use by boys and girls. *It is very explosive.* Hydrogen is colorless, odorless and tasteless and the lightest substance known, being 14.5 times as light as air, 11,160 times as light as water, and 151,700 times as light as the metal mercury. It enters into the composition of all plants and animals. This element constitutes a large part of our coal, wood and petroleum. But the great storehouse of hydrogen in the world is in the vast amounts of water which occur in nature. Hydrogen occurs in the free state in the gases from some volcanoes, in many natural gas wells, and in the atmosphere of the sun and of some of the fixed stars.

Because of its lightness, hydrogen was formerly used for filling airships, dirigibles and balloons (at left), but has now been largely replaced by helium. During World War I and World War II hydrogen was used industrially in enormous quantities, and generated cheaply by treating the metal aluminum with sodium hydroxide. These materials were economical to use in large quantities and could be easily transported from place to place.

HELIUM AND AVIATION

Owing to the explosive nature of hydrogen gas it is not used today for inflation of dirigibles and balloons. The gas helium, which is non-combustible, has been substituted for this purpose. The disastrous loss of many dirigibles has demonstrated the danger of using hydrogen as a bouyant gas. A leak in a gas bag is sufficient to cause the destruction of a dirigible. The United States is fortunate in having supplies of natural gas in the State of Texas and in other parts of this country which contain sufficient helium to allow its separation on a commercial scale. Helium, discovered previously in connection with the rare gases of the atmosphere, is surpassed only by hydrogen in lightness but it is superior to hydrogen as it is non-inflammable.

The densities of hydrogen, helium, oxygen and nitrogen, as compared to air, are recorded in the table below. Notice that helium is practically twice as heavy as hydrogen and consequently has less bouyant power.

Gas	Specific Gravity at 0°C and 760 mm.
Air	1.000
Hydrogen	0.06952
Helium	0.13804
Nitrogen	0.96724
Oxygen	1.10527

REDUCTION

Hydrogen can be heated to a high temperature in the absence of oxygen without danger. It interacts with many substances under specific conditions resulting in removal of oxygen and production of substances richer in hydrogen. Oxygen need not necessarily be removed by hydrogen to accomplish a reduction, but hydrogen may add or combine directly leading to reduction of the substance. In other words, to *reduce* is the reverse of to oxidize, and the process of removing oxygen from a substance is called reduction. We do not necessarily need to use hydrogen to accomplish reduction any more than we need to utilize oxygen to accomplish oxidation. Hydrogen peroxide and nitric acid for example are chemical reagents which react as oxidizing agents.

We likewise have chemical reagents which have the power to bring about reduction. Such substances are called reducing agents. Hydrochloric acid and sodium bisulphite are examples of well known inorganic reducing agents.

BLEACHING

Bleaching is a commercial term widely used in the industrial textile field. It means to decolorize. Cloth is bleached by exposure to sunlight and also by the action of chemicals to remove all color, and produce white goods. Both oxidizing and reducing agents can be utilized to bleach fabrics. Hydrogen peroxide is an oxidizing agent which finds wide application as a bleaching agent; for example, to bleach silk. The gas, sulphur dioxide, is a cheap commercial reducing reagent which is a valuable bleaching agent and used in industry, especially in the millinery and straw hat trade.

NITROGEN

Nitrogen is a colorless, tasteless gas. It forms four-fifths of the bulk of air. Chemically speaking, it is very inactive, and if it were not for its presence in the air, life on the earth would be destroyed. This gas dilutes the oxygen of the air, thereby preventing destruction by oxidation of living and inanimate material. Nitrogen is slightly lighter than air (see density value on page 33 in this series). Since nitrogen is a very inactive element, it combines with few other elements. However, those elements with which it combines form very interesting classes of compounds. Many of these compounds containing nitrogen occupy an outstanding position in chemical industry. The fact that nitrogen is a very inactive element would lead one to think that its compounds would be unstable and would decompose easily. This is true, and most of the high explosives used for military operations are nitrogen compounds. Many of the most highly explosive substances are compounds of nitrogen with hydrogen, oxygen and sulphur. The force of an explosion is due to the tremendous volume of gases that are simultaneously formed when an explosive compound is detonated in a cannon or a shell. Gunpowder, nitroglycerine, nitrocellulose, or gun cotton, picric acid, trinitrotoluene (T.N.T.) and tetranitroaniline (T.N.A.) and ammonium nitrate are some very common high explosives of nitrogen employed in military operations. A peace time use of nitrogen is in explosives as in the form of dynamite and nitroglycerine.

Much can be said about nitrogen and its services both in peace and war. As a necessary constituent of protein, this element stands among the first of the 92 elements in every-day importance. Because nitrogen is needed by the animal body, there must be a constant source of supply of this element for growing plants, since animals obtain their nitrogen supply from plants or from other animals which have in turn received nitrogen from plants. Plants obtain their nitrogen from the soil, having been built up by the action of bacteria which are capable of fixing nitrogen of the air, or by the addition of fertilizers containing nitrogen.

Nitrogen also is used in many other industries, among them lacquer, coated textiles and plastics.

Nitrogen combines with oxygen to form several oxides. Nitric acid (HNO_3) is a compound of nitrogen with hydrogen and oxygen. This important acid is really the basis of most explosives containing nitrogen. It is also used in the manufacture of such commercial products as dyes and fertilizers. Nitric acid is manufactured on a commercial scale by treating Chile saltpetre with sulphuric acid. This acid, and also ammonia, are manufactured today in enormous quantities from nitrogen of the atmosphere, by the "fixation of nitrogen" process.

EXPERIMENT 33—Preparation of Nitric Acid

Equipment: Potassium nitrate, sodium bisulphate, test tube, measure, candle or alcohol lamp, blue litmus paper.

Operation: Put 4 measures of potassium nitrate and 4 measures of sodium bisulphate in a test tube. Add 4 or 5 drops of water. Moisten a piece of blue litmus paper and place it over the mouth of the tube. Now heat the test tube slowly over a flame, and notice that the blue litmus paper turns red. Remove the test tube from the flame and smell cautiously the fumes that are given off. These are nitric acid fumes and if they are led into water they will dissolve immediately, resulting in nitric acid solution.

Result: Potassium nitrate is a salt of nitric acid and when it reacts with the sodium bisulphate the free nitric acid is liberated.

EXPERIMENT 34—Nitrogen from the Air

Operation: Repeat Experiment 29 and remove all oxygen from air with a burning candle. Remove the jar from the pan by placing a small glass plate or a piece of glazed cardboard over the mouth of the jar while under water. Set the bottle right side up, keeping the cover on so as not to lose any of the nitrogen. Light a splinter of wood and removing the cover from the mouth of the jar containing nitrogen, plunge the splinter into the jar.

Result: Notice that the flame goes out, showing that nitrogen does not support combustion. Perform a similar experiment by plunging a burning splinter into a jar of ordinary air. Observe the difference in behavior. Ordinary air contains sufficient oxygen to support combustion.

AMMONIA

Every boy and girl is more or less familiar with the sharp and characteristic odor of ammonia. This chemical serves a great many practical purposes. In the household, at the kitchen sink or in the bathroom, it is used for cleaning purposes and for softening water. Ammonium carbonate is a salt which readily gives up its ammonia fumes under ordinary temperatures and is commonly used as smelling or aromatic salts. Ammonia gas is very soluble in water, and the acqueous solution-ammonia water—is the form in which it is commonly sold in stores. Ammonia is condensed and shipped in enormous quantities today in steel cylinders. Ammonia in this form finds a wide use in artificial refrigeration. Formerly most of the industrial ammonia was obtained from the destructive distillation of coal. It is a valuable by-product in the manufacture of illuminating gas. Today the gas is made in large quantities synthetically by combining nitrogen of the air with hydrogen gas.

EXPERIMENT 35—Preparation of Ammonia

Equipment: Sodium carbonate, ammonium chloride, measure, test tube, candle or alcohol lamp, red litmus paper.

Operation: Put 2 measures of sodium carbonate and 2 measures of ammonium chloride in a test tube, add 1 spoonful of water and heat gently. Remove the test tube from the flame and notice the smell of ammonia in the mouth of the test tube. Place a strip of moistened red litmus paper over the mouth of the test tube and notice the change of color in the paper. *It will turn blue* proving the presence of an alkali.

Result: Any ammonium salt when heated in the presence of a base or alkali (lime water, sodium carbonate, or sodium hydroxide) will give off free ammonia gas and this reaction is used as a test for the ammonium group (NH_4) in a compound. If the gas produced by heating the above mixture is passed into cold water, it is dissolved and ammonium hydroxide is formed.

EXPERIMENT 36—Formation of Ammonia by Decomposition of Organic Matter

Most organic matter contains nitrogen and when heated in the presence of an inorganic base, like calcium oxide, the nitrogen is partially liberated in the form of ammonia (NH_3). Ammonia was first prepared by this method and was called "Spirits of Hartshorn."

Equipment: Calcium oxide, wool, hair, silk, finger nail clippings, measure, test tube, test tube holder, red litmus paper and alcohol lamp or candle.

Operation: Place a small quantity of wool, hair, silk or finger nail clippings in a test tube, and add 3 measures of calcium oxide and 5 or 6 drops of water. Place a small strip of moistened red litmus paper over the mouth of the tube and gently heat the tube over a flame. Notice that the red litmus paper turns blue, showing that a volatile base is formed. Remove the tube from the flame and smell at the mouth of the tube. You will recognize the odor of ammonia if the experiment is properly conducted.

Result: Proteins are nitrogen compounds and are very stable combinations. Many proteins, however, liberate ammonia gas when heated at a high temperature with calcium oxide or unslaked lime.

EXPERIMENT 37—Dissociation and Vaporization of an Ammonium Salt

Equipment: Ammonium chloride, measure, test tube, test tube holder, alcohol lamp or candle.

Operation: Put one measure of ammonium chloride in a clean dry test tube and heat slowly over a flame. Notice that the dry salt (NH_4Cl) passes into the vapor state and then condenses again on reaching the cooler part of the tube to form solid ammonium chloride.

Result: What really takes place when heat is applied is that the ammonium chloride is split up into molecules of ammonia (NH_3) and hydrochloric acid (HCl) which re-combine in the cool part of the tube to form again solid ammonium chloride.

Explanation: Reaction of base and acid combine to form a salt.

The chemical formula for ammonia is NH_3. The chemical formula for hydrochloric acid is (HCl).

Here we are dealing with a *base* and an *acid.* A base and acid cannot exist in the presence of each other. For example, ammonia (NH_3) and hydrochloric acid (HCl) combine immediately and lose their characteristic chemical properties. They interact to form a neutral salt or ammonium chloride (sal ammoniac) as is expressed in the following chemical equation:

$$NH_3 + HCl = NH_4Cl$$

(Ammonia) (Hydrochloric Acid)

(Ammonium chloride)

FIXATION OF ATMOSPHERIC NITROGEN

The name given to the commercial process of fixing or utilizing the free nitrogen occurring in the air to make such chemical compounds as nitric acid (HNO_3), ammonia (NH_3) and nitrates, is *fixation of nitrogen.* The supply of natural nitrates of our earth, such as Chile saltpetre (Sodium Nitrate $NaNO_3$) is limited and the increased commercial demand for nitrogen and its useful compounds is so large today, that it has become necessary to draw on the enormous resources of the atmosphere surrounding the earth as a source of supply.

Over one-half of the world's requirements of nitrogen compounds (inorganic and organic) are supplied today by synthetic or artificial processes. These are (1) the electric process, (2) the synthetic ammonia process and (3) the various applications of the cyanamide process.

DISSOCIATION OF WATER INTO THE ELEMENTS OXYGEN AND HYDROGEN

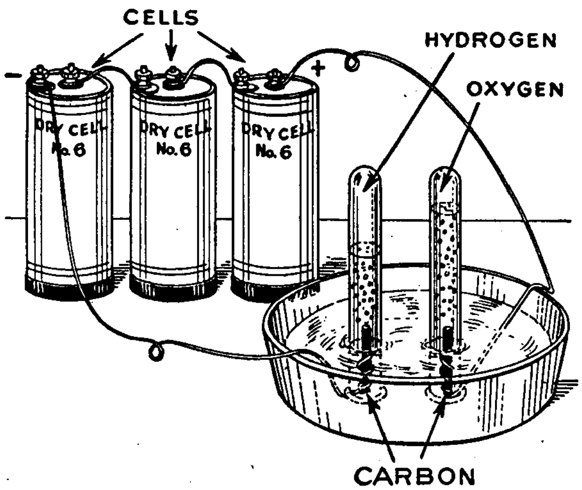
EXPERIMENT 38—Decomposition of Water by an Electric Current (Electrolysis)

It can be definitely proven that water is composed of two gases, hydrogen and oxygen.

Equipment: For this experiment you will need two or three good dry cells, a shallow dish, some insulated copper wire, two test tubes and two pieces of carbon rod or clean lead, and sodium bisuphate.

The apparatus is set up as shown in the illustration (Page 28). The cells are connected together in series with copper wire by joining the carbon binding post of one cell to the zinc binding post of the other cell.

To the free zinc and carbon posts of the end cells attach the long pieces of copper wire and fasten the short pieces of carbon rod or lead to the free ends of these wires. It is important to scrape the free ends of the wire with a knife to insure a clean surface.



Operation: Now fill the dish about half full of a solution of sodium bisulphate, using one teaspoonful of the compound to each glass of water. Fill two test tubes with this solution, close the mouth of each with your thumb and invert it upside down in the dish, being careful not to allow any bubbles of air to get into the test tubes.

Now put one of the carbon or lead electrodes under each test tube, being careful not to allow any air to get into the tubes.

Result: Notice that immediately bubbles of gas begin to rise in each test tube from the carbon or lead electrodes and gradually force the solution down out of the mouth of the tubes. Also notice that there is twice as much gas formed in one tube as in the other. This is because water is composed of two volumes of hydrogen to one volume of oxygen. The electrode at which the hydrogen is liberated is the negative (zinc) electrode or cathode, and the electrode at which the oxygen is liberated is the positive (carbon) electrode or anode.

Note: Test the gases collected in the two test tubes for hydrogen and oxygen respectively according to directions already described in your chemistry manual.

EXPLANATION OF THE ELECTROLYSIS EXPERIMENT

Hydrogen, when in solution as an *ion*, has a positive charge of electricity. When an electric current is passed through the solution, the positive hydrogen is attracted to the negative electrode or cathode where it loses its charge of electricity and is given off as a gas.

Oxygen, on the other hand, is formed in the solution through chemical reaction, and is liberated at the positive electrode or anode. Being very slightly soluble in water, it is also given off as a gas.

Sodium bisulphate is added to make the water a conductor of electricity, since pure water itself is a very poor conductor.

WATER

At ordinary temperatures pure water is a tasteless, odorless, transparent liquid; colorless in thin layers, but distinctly blue when seen in large masses. It is about 773 times heavier than air.

Water consists of two elements—both gases—hydrogen and oxygen, and they occur in the proportion of 89 per cent of hydrogen by weight. By volume, water consists of two parts of hydrogen to one part of oxygen.

Water occurs very abundantly throughout the earth. Vast areas of the colder regions of the globe are covered with it in the form of ice. In the liquid state water covers about five-sevenths of the earth's surface, reaching in some places a depth of nearly six miles. Large quantities occur in the soil, and as a vapor it is an essential constituent of the atmosphere. More than half the weight of living organisms consist of water. A great many substances dissolve in water, so that water is known as a very good solvent. A substance is said to be dissolved when none of the particles of the substance can be seen in the liquid nor separated from the liquid by filtering. A dissolved substance can usually be recovered from a liquid by evaporating the liquid.

The importance of water in the growth of plants can be judged by the fact that from 30 to 120 gallons of water are required by an ordinary plant for the production of each pound of dry substance present. The greater part of this water which is taken in by the roots is given off by the leaves. The flow thus maintained serves to keep the plant cells fully distended thereby preserving the form of the plant, and enabling it to carry on its vital processes. Even though the largest proportion of the water taken in during the growing life of a plant is given off through its leaves, a considerable amount is retained. In succulent plants, about 90 per cent of the complete weight of the plant is water, this water being an actual part of the organism.

PURITY OF WATER

Water which falls on our earth as rain, snow and general precipitation is really the purest form of water to be found in nature. But this water is not absolutely pure from a chemical point of view. The impurities

Table on HARD WATER and SOFT WATER

	Hard Water	Soft Water
With soap	Forms soap curds, and makes cleaning difficult.	Produces luxurious lather easily.
In bathing	Makes skin drawn, clogs the pores, preventing proper elimination of poisons by the skin.	Keeps skin soft and clean. Keeps pores free.
In shampooing	Deposits soap curds in hair, dulling lustre.	Makes hair clean and glistening.
In beauty culture	Impossible to thoroughly cleanse skin.	Keeps skin clean and free from blemishes.
In shaving	Makes a water-proof coating of soap curds around each whisker, causes razor blade to be dulled.	Thoroughly softens the toughest beard. Makes shaving easy—blades last longer.
In washing dishes	Causes streaks on china and glassware.	Dishes can just be drained clean. Many pieces do not need to be touched by a towel.
In removing grease	Makes work difficult.	Cuts grease like magic.
In home laundering	Wastes soap, discolors clothes, often necessitates extra hard washing.	Saves 50-80% of soap formerly used. Makes clothes soft, sweet, clean. Prolongs their life.
In washing woodwork, porcelain, etc.	Calls for hard scouring—harms finish.	Preserves lustre.
In pipes	Scale chokes down flow of water. Makes faucets drip.	Keeps pipes free from scales. Faucets seat tightly.
In water heater	Deposits rock-like scales in coil.	Coils last longer. Minimum fuel consumed.
In steam boilers	Wastes fuel. Steam slow to come up.	Saves fuel. Steams quickly.
In automobile radiators	Will eventually clog fine tubes, causes overheating, boiling away of water.	Keeps radiator in efficient condition. Keeps motor operating at proper temperature.
In making pie crust	Tends to make crust tough.	Makes crust flaky and tender.
In cooking green vegetables	Toughens them, causes loss of flavor.	Keeps them tender with all natural freshness. Needs less cooking.
In making tea and coffee	Impairs flavor.	Makes them taste better.
For drinking	Hard water is unpleasantly flavored.	A water softener like permutet removes calcium and magnesium.
For care of baby	Causes rashes blamed on prickly heat. Milk bottles get scummy—harder to clean.	Eliminates many skin troubles. Makes bottles glisten.
For hygienic cleanliness	Makes it hard to remove dirt and foreign matter. Prevents thorough cleaning.	Cleans thoroughly and easily. Insures absolute cleanliness.

found here are due to the presence of fine dust particles of different composition (dirt), and small traces of mineral and organic matter.

WATER BY DISTILLATION

In the chemical laboratory the purest form of water is obtained by distillation. Distilled water prepared in this way is used in automobile and airplane storage batteries. Storage cell water should be preserved in clean glass bottles.

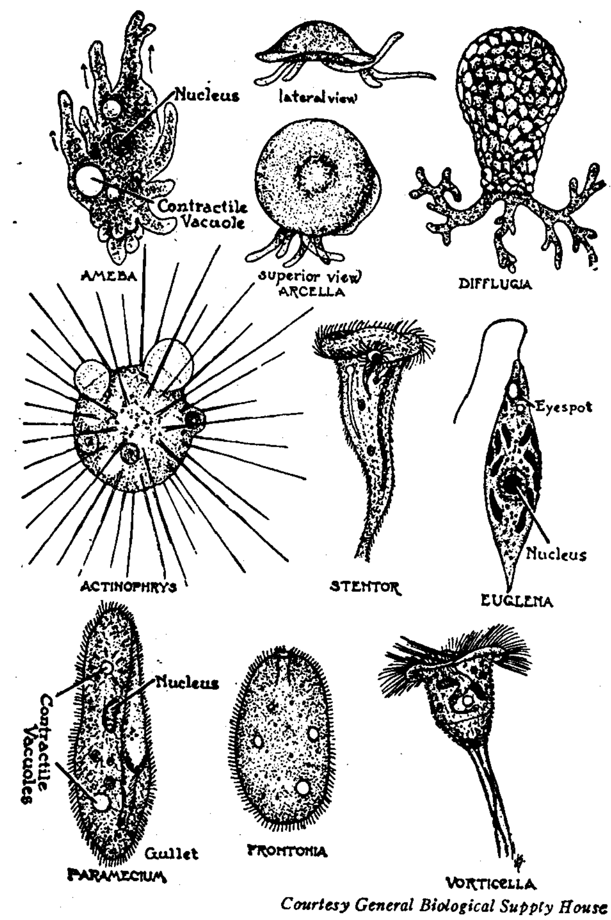
AQUEOUS SOLUTION

EXPERIMENT 39—A Soluble Substance

Dissolve one or two teaspoonfuls of common salt in a glass of water. Notice that you can no longer see the salt. If this liquid was filtered, you would find that nothing would remain on the filter paper.

Pour the solution into a clean pan and heat on a stove until all the water is driven off. Notice that a white solid remains. Allow this to cool and taste a little of the solid. It is the same salt you dissolved in the water, proving that the salt when in solution had undergone a physical change only.

Forms of animal and plant life often found in water → as seen through a powerful microscope.



Courtesy General Biological Supply House

EXPERIMENT 40—Separation of Starch from Sugar

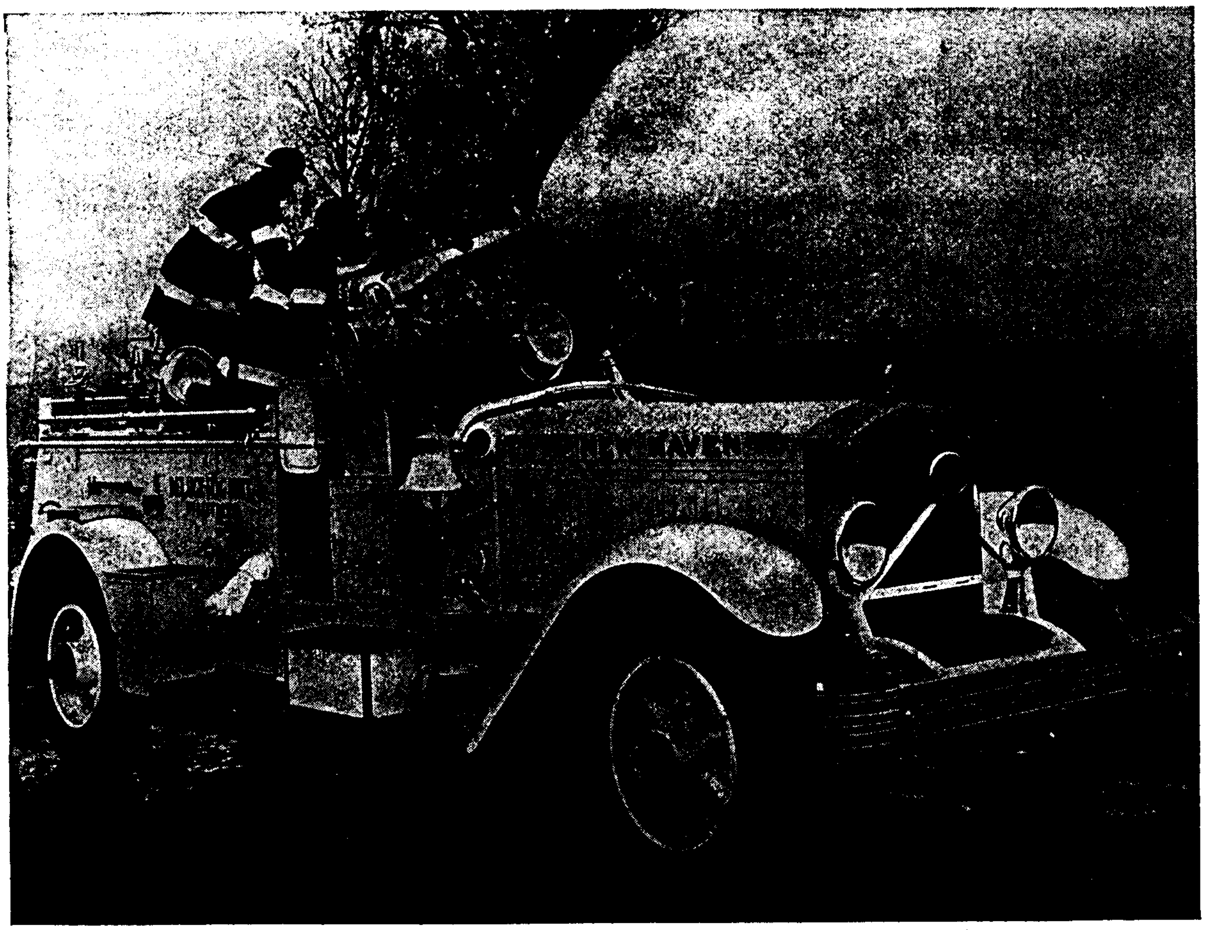
Equipment: Sugar, starch, cellophane bag, fruit jar.

Operation: Prepare a mixed solution of sugar and starch and place in a cellophane or parchment bag. Suspend in a jar of water and sugar will diffuse through the membrane into the outside liquid. Taste the water after a few hours and you can detect the sugar by the taste. Starch *will not* diffuse through the parchment into the outside solution. To prove this, test some of the outside solution with starch-iodide paper. If starch is present a blue coloration will be produced.

Result: By changing the outside water several times all the sugar can be separated from the starch in the cellophane or parchment bag.

EXPERIMENT 41—Separating Sugar from a Protein

Repeat the above experiment using a solution of egg white and sugar. Sugar will diffuse through the membrane. Like starch, the egg protein *will not diffuse* through the membrane.



Courtesy Fire Dept., New Haven, Conn.

CHEMISTRY FIGHTS FIRE

This Water-Fog Unit is most useful in fighting oil fires. Water is broken up into a spray of fine particles which readily absorb heat and quickly reduce temperature of burning material to below kindling or flash point, stopping combustion.

For liquids having a flash point above room temperature, such as gasoline, carbon dioxide is more effective. It chokes out fire by displacing oxygen which is necessary to combustion.

PART VI

CHEMICAL ACTION IN WATER SOLUTION

MOST chemicals, as a rule, do not interact when brought together in the dry form. Their molecules are inert and need to be activated before a chemical change can take place. When water is present so that solution can take place they then react readily. By solution in water the molecules are changed or activated, and the particles of the compounds are brought much closer together. Other solvents besides water can be used to promote chemical change. For example, liquid ammonia can serve as a solvent for carrying out many interesting reactions that cannot be produced in water solution. Alcohols, acetone, benzene and other organic reagents are used in place of water as solvents for carrying out organic reactions. The organic chemist and inorganic chemist apply an entirely different technique in accomplishing chemical reactions.

EXPERIMENT 42—A Color Change due to Solution

Equipment: Tannic acid (purchase separately), ferric ammonium sulphate, measure, test tube.

Operation: Make a mixture of one measure of tannic acid and one measure of ferric ammonium sulphate on a piece of paper and notice that there is no evidence of reaction. Transfer this mixture to a clean, dry test tube and fill half full of water.

Result: Observe the formation of a black colored product showing that there was a chemical reaction due to the presence of water. The black product formed is a basic constituent of many inks.

DRY ICE EXPERIMENTS

Dry ice is a commercial term assigned to solid carbon dioxide gas (CO_2). This solid material slowly reverts to carbon dioxide gas on exposure to the atmosphere. The student should note the difference between ordinary ice and the term—*dry ice.*

EXPERIMENT 43—Extinguishing a Burning Candle

Equipment: Dry ice, candle, fruit jar.

Operation: Place a burning candle in the bottom of a pint fruit jar. Then drop a spoonful of dry ice to the bottom of the jar.

Result: Notice what happens when the dry ice turns to gas. A candle cannot burn in carbon dioxide gas. The gas does not support combustion.

EXPERIMENT 44—Burning Sulphur

Repeat the preceding experiment, substituting burning sulphur for the candle.

EXPERIMENT 45—Floating a Sunken Ship

Equipment: Dry ice, rubber bag, jar, spoon.

Operation: Place a spoonful of dry ice in a thin rubber bag such as a small toy balloon and tie with a silk thread so that gas will not escape. Then attach a weight to the bag so that it will sink to the bottom of the jar of water.

Result: Note the result as the solid dry ice gasifies inside the rubber bag. The bag will finally float. How do you explain this?

EXPERIMENT 46—Measuring Dry Ice Pressure

Equipment: Mercury, dry ice, u-tube.

Operation: Take a U-tube made of glass tubing about 5 mm. diameter and fill the lower curviture with a little mercury. For information on making a U-tube, refer to the section in your manual on glass blowing. Then drop some dry ice into one arm of the tube and close the end of the U-tube immediately with a tight rubber stopper.

Result: Observe what happens to the mercury as the dry ice vaporizes in the tubing. It will be forced by gas pressure into the second arm of the tube and finally be expelled if the arm is of small diameter.

EXPERIMENT 47—A Miniature Gas Volcano

Equipment: Dry ice, salt seller, jar.

Operation: Take a heavy-walled salt seller and fill it with dry ice, then place in upright position on the bottom of a tall glass jar filled with water.

Result: As the dry ice vaporizes the gas will rush through the openings of the salt cellar lid and the bubbles will travel to the surface of the water. Hold an electric bulb back of the jar of water during the bubbling.

ARTIFICIAL REFRIGERATION

The application of the so-called negative heat of solution of common salt is made use of in the ice-salt freezing mixture so commonly used for making ice cream. Common salt, just like ammonium chloride, lowers the temperature of water when dissolved in it (endothermic change). Therefore, when salt is added to a mixture of ice and water whose temperature is just at the freezing point, the salt dissolves and in doing so, lowers the temperature of the solution several degrees below freezing, thereby affording us a very convenient freezing mixture.

EXPERIMENT 48—How to Make a Freezing Mixture

Equipment: Ice, test tube, glass, salt.

Operation: Mix together a glass full of cracked ice and one-half glass full of salt. Try the effect of this freezing mixture upon water by placing a test tube half full of water in it and allowing to stand for several minutes.

Result: Notice that the water in the test tube will freeze to solid ice after a short time. Use your thermometer for observing the changes in temperature in all these experiments.

CHANGING THE FREEZING AND BOILING POINTS

It is easily demonstrated that when a substance is dissolved in water the freezing point of the resulting solution is lower than the normal freezing point of water. This is the reason why salt is sometimes thrown on slippery sidewalks. It melts the ice by lowering the freezing point of the water.

It can also be shown that when a substance is dissolved in water, the boiling point of the resulting solution is higher than that of pure water.

EXPERIMENT 49—Solution Lowers the Freezing Point of Water

Operation: Make a freezing mixture as shown in the preceding experiment. Now add a spoonful of common salt to a test tube half full of water and shake until dissolved. Place this test tube of salt solution with another test tube half full of water in the freezing mixture and allow to stand. Note the change in temperature by means of a thermometer.

Result: Notice that the water freezes but that the salt solution does not, thereby proving that solution lowers the freezing point of water.

EXPERIMENT 50—Solution Raises the Boiling Point of Water

Equipment: Salt, test tubes, test tube holder, alcohol lamp or candle, thermometer.

Operation: Make a salt solution by dissolving a spoonful of salt in a test tube half full of water. Hold this solution, together with a test tube half full of water, over flame, giving each about the same amount of heat.

Result: Notice that the water will boil before the salt solution does, thereby proving that solution raises the boiling point of water. Note the change in temperature by means of a thermometer.

Note in performing Experiment No. 49 and No. 51, an ordinary household thermometer may be used.

SUPERCOOLING

It is possible to cool a liquid below its freezing point. Water, for example, can be cooled below its freezing point, 32 degrees Fahrenheit or 0 degrees Centigrade and still remain a liquid. When in this state, water is said to be undercooled.

EXPERIMENT 51—Undercooled Water

Equipment: Ice, salt, test tube, glass, thermometer.

Operation: Make a freezing mixture as explained in previous experiments. Place in the freezing mixture a test tube one-third full of water and keep the test tube quiet.

Result: The temperature of the water in the tube may go down as far as 8 to 10 degrees below 0 degrees Centigrade or between 18 and 14 degrees Fahrenheit, and the water still remains in the liquid form. Read the temperature of your solution by means of a thermometer. If a small crystal of ice is now dropped into the test tube or the water in the test tube stirred, it will immediately freeze, the temperature then rising to the freezing point 0 degrees Centigrade or 32 degrees Fahrenheit. Follow the changes in temperature with your thermometer.

DEGREE OF SOLUBILITY IN WATER

Most substances dissolve more readily in hot water than in cold water. There are a few exceptions, however, calcium hydroxide being a good example of a substance which is more soluble in cold water than in hot water.

EXPERIMENT 52—Effect of Temperature in Solubility

Equipment: Nickel ammonium sulphate, measure, test tube, alcohol lamp or candle.

Operation: Put 7 measures of nickel ammonium sulphate in a test tube ¼ full of water and shake well. Notice that some of the solid remains undissolved. Now heat the tube slowly and notice that all the solid goes into solution, showing that some substances are more soluble in hot water than in cold water.

Result: Allow the test tube to cool undisturbed and notice the beautiful green crystals of nickel ammonium sulphate that separate out on cooling.

EXPERIMENT 53—Temperature Effect on Solubility

Equipment: Calcium oxide, measure, test tube, alcohol lamp or candle, test tube holder.

Operation: Add one measure of calcium oxide to a test tube full of water. Shake several times and allow the tube to stand until the liquid becomes clear. Pour some of this clear liquid into another test tube and heat slowly over a flame.

Result: Notice that the liquid becomes cloudy or turbid, proving that calcium hydroxide, formed when calcium oxide was added to the water, is less soluble in hot water than in cold water.

EXPERIMENT 54—Diffusion

Equipment: Mixed dyes, glass.

Operation: Fill a clean glass nearly full of clear water and let it stand for a minute or two to become quiet. Now add a small quantity of mixed dyes and watch closely what occurs.

Result: As the substance dissolves the color seems to flow out of the small crystals, gradually spreading over the bottom of the glass. After a few minutes the color will begin to diffuse upward through the liquid until finally after several hours the entire solution will be a uniform color.

EXPERIMENT 55—Testing for Acidity

Operation: Moisten a piece of blue litmus paper with a drop of the water you are testing.

Result: If the blue litmus paper turns red or pink, the water is slightly acid.

EXPERIMENT 56—Testing for Lime

Operation: To a test tube of water add two measures of sodium carbonate and shake until dissolved.

Result: If the water, after standing a while, shows a white turbidity it contains a considerable portion of lime.

EXPERIMENT 57—Testing for Carbon Dioxide

Operation: Make up a solution of lime water by adding half a measure of calcium oxide to half a test tube of water. Shake well and let settle for a few minutes. Add a few drops of this clear lime water to a test tube of water which is to be tested.

Result: If a white turbidity is formed the water contains carbon dioxide.

Explanation: The presence of carbon dioxide in water causes it to effervesce and gives it a sparkling test. Soda water and many bottled mineral waters contain carbon dioxide.

EXPERIMENT 58—Testing for Sulphur

Operation: Sulphur is present in some mineral waters in the form of hydrogen sulphide. It may be detected even if in very small quantities by dropping in the water a small piece of sulphide test paper.

Result: If sulphides are present the paper will turn black or brown.

EXPERIMENT 59—The Boiling Point of Water, Compared with that of Salt Water

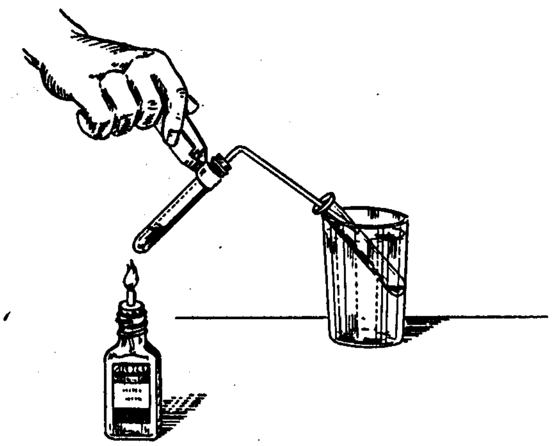
Operation: Dissolve 10 measures of common salt in a test tube half full of water and fill another test tube half full of plain water. Now heat these two test tubes over a candle or alcohol lamp flame. Hold them so that each tube will get just the same amount of heat.

Result: You will find that the water will boil before the salt solution does. Hold the tubes carefully so the boiling water will not spurt out on you.

EXPERIMENT 60—Purification of Water by Distillation

Equipment: Copper sulphate, measure, test tubes, gas delivery tube, candle or alcohol lamp.

Operation: To a test tube half full of water add one measure of copper sulphate. Shake well until the salt is dissolved. Then attach the gas delivery tube by means of the perforated cork to the mouth of the tube and insert the end of the delivery tube into a clean test tube which is immersed in a glass of cold water.



Now heat the test tube containing the copper sulphate solution and boil the liquid for several minutes. You will notice that clear, colorless water condenses in the test tube immersed in the glass of water.

Result: What you really did was to distill the copper sulphate solution. That is, the water simply passed over on heating in the form of steam which condensed to water on passing into the second cooler tube. The copper sulphate being nonvolatile, remained behind so that we could remove all the copper sulphate from the water in this way.

WATER OF CRYSTALLIZATION

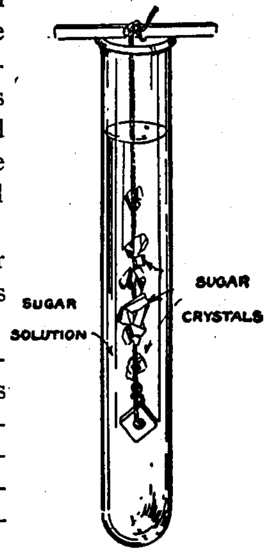
Many compounds contain chemically combined water. Water occurring in compounds in this way is known as water of crystallization. Ferrous sulphate, nickel sulphate, and copper sulphate, for example, contain water of crystallization.

Some substances give up or lose their water of crystallization by simple exposure to air. Such substances are called efflorescent substances and a good example is sodium sulphate.

On the other hand, certain substances on exposure to the air take up water from the air and in some cases dissolve in this water to form a liquid. Substances of this class are called deliquescent substances of which a good example is calcium chloride.

EXPERIMENT 61—Crystal Formation (Rock Candy)

Operation: Dissolve as much sugar as possible in a test tube half full of boiling water. Suspend a thread or string in this solution by hanging a small weight on the end and allow the contents of the tube to cool slowly undisturbed.



Result: After a time the sugar will appear in large crystals upon the string.

Large crystals of many compounds can be formed in this manner, or by allowing a solution to cool slowly in a crystallizing dish. Slow cooling is essential if good crystalline development is desired.

TESTING WATER

Absolutely pure water is never found in nature. The impurities found in water are of two classes. The inorganic, or those that come from the rocks; and the organic, or those that are formed from the decay of animal or vegetable substances.

The principle inorganic matter found in water is common salt and compounds of calcium, magnesium and iron. Waters containing such substances in solution are commonly spoken of as hard waters, or, if large amounts of mineral matter are present as mineral water. Some of the natural mineral waters possess valuable medicinal properties and consequently are set aside and protected for public use. Springs of this type are located in several sections of the United States. The salts occurring in hard waters do not injure the water for drinking purposes but they form insoluble compounds with soap so that we cannot wash with them.

In addition to mineral matter, natural waters contain more or less organic matter in solution or held in suspension. This organic matter is not necessarily harmful but quite often this is accompanied by certain forms of micro-organisms or living bacteria which may be very injurious to life. Typhoid fever may be contracted from drinking water containing bacteria of this kind. Bacteria when found in drinking water is generally destroyed by adding bleaching powder or chloride of lime to the water. Chlorine gas and ozone are also used for the same purpose with good effect.

SEWAGE CONTAMINATION

Water is made unfit for drinking purposes by contamination with sewage waste. Therefore, it is very important to make bacteriological tests of drinking water occasionally to avoid the possible spreading of disease. In fact, the sewage disposal problem of today is one that is coming to the front because sewage improperly treated is a menace to public health. The unthinking citizen believes the problem of sewage disposal is solved when the toilet is flushed or the bath tub drained. Actually, the problem may be said to commence at this point. Many cities discharge their sewage untreated into rivers, but seldom do we find any stream in sufficient volume or speedy enough flow to make such a method efficient. Sewage allowed to stand in rivers leads to unsightly appearances and very offensive odors, and such water is unfit for drinking purposes.

BACTERIA IN INDUSTRY

Many diseases like typhoid fever result from bacterial infection, but it should be emphasized here that, contrary to a popular belief, all bacteria are not harmful to man. Great prominence has been given them as causative agents in disease. Therefore, it is perfectly natural that bacteria in general would be looked upon as organisms, which are harmful to man. This is not true. Taking the entire group as a whole, they are beneficial to man in that the good they do far out-weighs the harm. While we could suggest the most important processes which depend on microbial activity, the greatest single service is no doubt the part they play in nature in causing the decay of plant and animal bodies. Were it not for this process of decay, much of the supply of certain essential elements—elements which we could not do without—would remain locked up in the dead bodies of plants and animals so that what remained would not be sufficient for the growth and development of living plants and animals. Let us then bear in mind that bacteria are not, as generally supposed, undesirable and destructive, but like all other living objects, there are good ones and bad ones, and great as is the loss of life and property and suffering for which some bacteria are responsible, the beneficial effects out-weigh the harmful.

HARD WATER

When water does not lather well with soap, it is commonly known as hard water. The chief elements which are productive of water hardness are the alkali earth metals—calcium, magnesium, barium, and strontium. The most widely occurring of these is calcium, and when we are troubled with hard water we always look first for calcium contamination in some form. There are two kinds of hard water—temporary and permanent.

EXPERIMENT 62—Temporary Hardness—How to Get Rid of It

Operation: If you are able to obtain some hard water in your locality, test a half test tube full of it for temporary hardness by boiling for 3 or 4 minutes over a flame. If the water after boiling becomes turbid, that is, takes on a white milky color, it possesses temporary hardness.

Explanation: Temporary hardness is due to the presence of calcium bicarbonate, which is formed by the action of limestone (calcium carbonate) on the carbon dioxide dissolved in rain water. This form of hardness is easily gotten rid of by boiling. The heat drives off the excess of carbon dioxide and the calcium carbonate precipitates, giving the water a turbid or milky appearance.

It is due to this precipitate that kettles and boilers become gradually covered inside with a brown deposit. In large boiler pipes the deposit is known as boiler scale and constitutes a serious problem in manufacturing plants. On a large scale, slaked lime (calcium hydroxide) made by adding water to lime is used to soften water. The soluble bicarbonate is converted into the insoluble carbonate which separates out and is removed by filtering.

Try the following experiments on any samples of water that you may obtain:

EXPERIMENT 63—Testing for Odor

Operation: Fill a test tube half full of the water to be tested, shake the tube well and then smell at the mouth of the tube. Gently heat the tube for a few seconds and smell again.

Result: Notice any increase in odor. Heating usually drives out any dissolved gases which may be in the water as impurities. If the water has any disagreeable odor it may be contaminated with sewage of some sort.

EXPERIMENT 64—A Test for Color and Clearness

Operation: Examine the sample of water as follows: Hold a test tube of the water in front of a white sheet of paper and notice whether the water is colored and cloudy. If the water is colored or cloudy it is contaminated with impurities.

Result: Quite often when water is drawn from a faucet it appears a little milky. This is due to the high pressure of the water in the pipes. When this is so, allow the water to stand for two or three minutes and you will notice that the water is now clear.

EXPERIMENT 65—Test for Solid Matter in Water

Operation: Pour one-half test tube full of water to be tested into an ordinary tea cup and allow it to evaporate slowly down to dryness on the stove.

Result: Look for a residue or solid matter in the cup after all the water is driven off. Most waters contain small amounts of mineral salts and sometimes organic matter which are tested for in this way.

EXPERIMENT 66—How to Test for Acidity in Water

Operation: Add a small piece of blue litmus paper to a test tube full of water.

Result: If the litmus paper turns pink the water is slightly acid.

EXPERIMENT 67—How to Test for Lime in Water

Operation: Add 1 or 2 measures of sodium carbonate to a test tube one-half full of the water to be tested and allow to stand for 10 to 20 minutes.

Result: If the water becomes cloudy or turbid, lime is present.

EXPERIMENT 68—How to Test for Carbon Dioxide in Water

Operation: Add a few drops of clear lime water to a test tube three-quarters full of water to be tested.

Result: A white precipitate or a milky color is a test for carbon dioxide.

Explanation: Lime water is made by adding 1 measure of calcium oxide to a test tube one-half full of water, shaking and allowing any solid material to settle. The clear liquid is lime water. Common soda water such as is served at a soda fountain is simply pure water which has been saturated under pressure with carbon dioxide. When the pressure is relieved from such water the carbon dioxide bubbles out, producing what is known as effervescence.

Although great skill and many years of experience are required to blow glass apparatus such as that illustrated above, you can learn the fundamental principles of glass blowing from your Gilbert Glass Blowing Manual.

Part VII

INORGANIC CHEMISTRY

SULPHUR

SULPHUR is a very important element commercially, and plays a very significant part in the physiological processes of animal life. It is a yellowish, tasteless solid and is practically odorless. The odor commonly ascribed to sulphur is not that of sulphur itself, but is due to sulphur dioxide when sulphur burns or undergoes oxidation.

In the free state sulphur occurs chiefly in volcanic regions. Large deposits are found in Italy, Sicily, China, Ireland and India. Important deposits are found in this country in Louisiana and California. In Louisiana the sulphur is melted under ground by means of superheated steam and forced out under pressure through pipes. Sulphur also occurs in many important ores as sulphides, for example, galena or lead sulphide; cinnabar or mercury sulphide; zinc blend or zinc sulphide; realgar or arsenic sulphide, and in pyrite (iron sulphide).

Sulphur is used extensively in the commercial manufacture of many substances, such as gunpowder, fireworks, matches, dyestuffs, medicinal products, drugs and fertilizers. Without doubt the most important chemical containing sulphur is sulphuric acid which is manufactured in enormous quantity. It occupies a key position in chemical industry and is utilized in hundreds of manufacturing operations. Sulphuric acid is consumed in enormous quantities in the vulcanization of rubber, in the bleaching industry, and in the manufacture of disinfectants, insecticides and dyestuffs. The great demand for insecticides by growers of fruit and truck garden products has led to the study of many materials to be used for plant protection. Today this represents an enormous industry and there is a constant search for effective chemical insecticides. Sulphuric acid is a valuable intermediate in their manufacture.

THE DIFFERENT PHYSICAL FORMS OF SULPHUR

Sulphur occurs in nature in three different physical forms as follows:

a. Rhombic Sulphur
b. Prismatic Sulphur (or monoclinic)
c. Amorphous Sulphur (or plastic)

These are known as the allotropic forms of this very common element and exhibit widely varying physical properties—color, shape or crystalline form, texture, solubility and stability. All of the physical forms, however, possess identical chemical properties.

Note: Collect small samples of the three allotropic forms of sulphur, and preserve them in specimen tubes for future study and examination.

EXPERIMENT 69—Solubility of Sulphur in Carbon Tetrachloride

Equipment: Carbon tetrachloride (purchase separately), test tube, sulphur, evaporating dish or watch glass.

Operation: Place a minute piece of pure sulphur in a dry test tube and cover with ½ inch of carbon tetrachloride. Shake the contents occasionally and note what happens to the sulphur. After the sulphur is dissolved, pour the contents of the test tube into your evaporating dish. Note the appearance of the sulphur when the carbon tetrachloride evaporates.

Result: Sulphur is insoluble in water, but is soluble in carbon-tetrachloride and also carbon bisulphide.

EXPERIMENT 70—Preparation of Plastic Sulphur

Equipment: Sulphur, heating spoon, alcohol lamp or candle, glass of water.

Operation: Heat 5 measures of sulphur in your heating spoon over a flame until the specimen of sulphur completely melts. Then pour this molten sulphur into a glass of cold water. Examine by touch and note its pliability. Set this plastic modification to one side and examine it from time to time and note the physical changes.

Result: When hot molten sulphur is suddenly cooled it assumes a plastic form resembling crude rubber which is not crystalline. This is not a stable form of sulphur, and on standing undergoes a slow transformation into hard, crystalline rhombic sulphur.

EXPERIMENT 71—Use of Sulphur for Casting or Moulding

Equipment: Amorphous sulphur, alcohol lamp or candle, spoon, coin.

Operation: Melt in your heating spoon 5 measures of amorphous sulphur. Do not let it take fire by overheating. Then pour the molten sulphur over a coin and when cold and completely solidified remove the coin.

Result: You will find a perfectly smooth impression of the coin cast in the sulphur. Plastic sulphur serves many uses as a strong cement.

EXPERIMENT 72—Behavior of Sulphur at Different Temperatures

Equipment: Sulphur, alcohol lamp or candle, measure.

Operation: Heat 10 measures of sulphur in a small, dry test tube. Apply the heat very slowly and notice the different changes. First, the sulphur melts to a light straw colored liquid. Pour a little of this liquid into a glass of water, and then continue the heating of the tube, and observe, second, the change of color to brownish-black and the liquid becoming almost solid. On further heating, third, this solid becomes liquid again. Pour this liquid sulphur into another glass of water.

Result: The sulphur obtained when the straw-colored liquid was poured into water is called rhombic sulphur, while that formed when the dark black liquid was poured into water is called plastic (or "elastic") sulphur. This substance becomes brittle on standing for a few days. Sulphur undergoes three distinct changes in heating and each change corresponds to a certain temperature.

EXPERIMENT 73—Preparation of Lime—Sulphur Solution

Equipment: Calcium oxide, sulphur, test tube, measure, alcohol lamp or candle.

Operation: Put one measure of calcium oxide and one measure of sulphur into a test tube one-third full of water. Heat the test tube over a flame and boil for several minutes. Notice the yellow-colored solution that is formed.

Result: This solution is known as lime-sulphur solution and is used on a large scale for spraying fruit trees and destroying fungi.

Explanation: The calcium oxide reacted with the sulphur to form calcium sulphide, which is soluble in the the water. Filter a part of the calcium sulphide solution, and add acetic acid to the clear fluid until the solution tests acid to blue litmus paper. There will be an immediate evolution of hydrogen sulphide, evidenced by the odor. Allow 1 or 2 drops of the calcium sulphide solution to fall on a polished silver coin, and let stand for a few minutes. Then wash the coin with water and notice that a black spot of silver sulphide is formed. Exposure of silverware to eggs will produce a similar discoloration due to the presence of sulphur in eggs.

EXPERIMENT 74—Sulphur Dioxide from Burning Sulphur

Equipment: Sulphur, measure, spoon, alcohol lamp.

Operation: Put 2 measures of sulphur in the spoon and heat over the flame.

Result: The sulphur will suddenly take fire and burn with a blue flame. The gas produced, having a suffocating odor, is sulphur dioxide, and is formed by the oxidation of sulphur when it burns in the air.

SULPHURIC ACID

Sulphuric acid is one of the most important acids. The processes involved in its manufacture are more complicated than those other acids because it must be built up from its elements. Sulphuric acid is made commercially by either one of two methods.

First, the "Lead Chamber Process" or the older method. In this process sulphur dioxide is prepared by burning sulphur, or an ore of sulphur such as iron sulphide or pyrite. The sulphur dioxide is then conducted into large lead chambers where it comes in contact with oxygen, oxides of nitrogen and steam. These react to form sulphuric acid which is concentrated by heating.

The second method, known as the "Contact Process" is the more recent method and gives an acid of much higher purity. Although a more expensive method, it is compensated for by the quality acid obtained. In this process sulphur dioxide is formed the same as in the lead chamber process. It is then passed through a tube heated to 400 degrees Centigrade containing a catalytic agent which makes the sulphur dioxide combine with more oxygen to form sulphur trioxide. This is passed into water forming concentrated sulphuric acid.

EXPERIMENT 75—Preparation of Sulphuric Acid

Equipment: Sulphur, potassium nitrate, measure, test tube, alcohol lamp or candle.

Operation: Mix together on a piece of paper ½ measure of sulphur and ½ measure of potassium nitrate. Put ¼ of this mixture—no more—in a clean, dry test tube and heat slowly over a flame. Notice the white fumes which are given off. These fumes are sulphur trioxide. After the fumes stop coming off, stop the heating and place your thumb over the mouth of the tube.

When tube is cold, fill ½ full of water and shake, holding your thumb over the mouth. Test liquid with blue litmus paper and notice that it turns red.

Result: Sulphur trioxide combined with the water to form sulphuric acid.

HYDROGEN SULPHIDE

Hydrogen sulphide (H_2S) and water (H_2O) are members of the same chemical family. Water is a neutral substance, while hydrogen sulphide is a weak acid. The salts of hydrogen sulphide are called sulphides. Some sulphides are very insoluble in water and for that reason are used in analytical work. Certain metals can be separated from each other by means of their sulphides.

EXPERIMENT 76—How to Make Hydrogen Sulphide

Equipment: Paraffin, sulphur, sulphide test paper, measure, test tube, alcohol lamp or candle.

Operation: Cut a piece of paraffin about the size of a pea from a candle or other wax and put it in a test tube. Add 2 measures of sulphur, place a piece of moistened sulphide test paper over the mouth of the tube and heat tube slowly. Notice that the test paper turns black. This is a test for hydrogen sulphide gas. The test paper contains lead acetate and reacts with hydrogen sulphide to form a black precipitate of lead sulphide.

Remove the tube from the flame and smell cautiously. Note the resemblance of the odor to that of rotten eggs. Hydrogen sulphide is given off from several organic compounds, for example, when cabbage is cooked.

Result: Hydrogen sulphide is inflammable and when burned the hydrogen combines with oxygen to form water while the sulphur combines with oxygen to form sulphur dioxide.

Reducing Action of Hydrogen Sulphide

You probably have noticed that some white paints turn black after exposure for a long time. This is because such paints contain lead, chiefly in the form of the pigment, lead carbonate. The lead reacts with traces of hydrogen sulphide gas in the air to form black lead sulphide. To prevent this discoloration, zinc oxide is used as a base pigment in place of lead carbonate. When zinc oxide reacts with hydrogen sulphide a white precipitate of zinc sulphide is formed.

EXPERIMENT 77—How to Restore the Color of White Paint

Obtain some pieces of wood once painted which have become dark by the action of hydrogen sulphide. Wash these pieces with a little hydrogen peroxide solution and notice that they become white again.

EXPERIMENT 78—Silver Sulphide

Place ½ measure of sulphur on a bright silver coin and wrap in several thicknesses of paper. After a few days you will find a black spot of silver sulphide on the coin where the sulphur was in contact with it.

EXPERIMENT 79—Sulphur in Rubber

Rubber contains sulphur used in its vulcanization. Wrap a rubber band around a silver coin and you will find that it will turn black after a few days, due to the formation of silver sulphide.

EXPERIMENT 80—Sulphur and Silver Coin

A silver coin is turned black in a few hours by a paste of mustard and water, as mustard contains sulphur. Eggs also contain sulphur, and this is the reason why silver spoons turn black after contact with eggs.

THE HALOGENS

The elements that go to make up the halogen family are fluorine, chlorine, bromine and iodine. They are known as the Halogens, meaning "salt producers." Resembling each other very much in chemical properties, they differ widely in physical properties. Fluorine is a colorless gas, chlorine is a greenish-yellow gas, bromine a brownish-red liquid and iodine a purplish-black solid.

The halogens are very active substances, so that they never occur in the free state in nature. Their compounds are very abundant—those of chlorine, bromine, and iodine occurring in sea water. The most common of these is sodium chloride or common table salt.

As already stated, the halogens are very active substances. They combine with metals like copper, sodium, potassium, gold, silver, platinum, etc. to form metallic salts and react with non-metals like sulphur, antimony and arsenic to form compounds. They also react with hydrogen to form the corresponding acids, namely, hydrofluoric, hydrochloric, hydrobromic and hydriodic acids. Of the halogens, fluorine is the most active and iodine the least active. All four of the halogens find wide commercial applications.

Fluorine, Chlorine, Bromine and Iodine

Fluorine is the most reactive of all the halogen elements, and is, in fact, the most active element of nature. Very few substances are so stable that they are not attacked and changed by its action. It is a pale yellow gas with an irritating odor, and is characterized by its extreme corrosive action upon the eyes and skin. The minerals, fluorspar and cryolite, are the commercial natural resources of this active element. They are used in the manufacture of aluminum and glass.

Chlorine is used extensively as a bleaching agent and germicide. Chlorine gas is shipped in bulk compressed in iron cylinders. It is widely used for water purification. It also comes on the market as bleaching powder or chloride of lime. The corresponding acid, hydrochloric acid, is an important technical acid and is used for a number of purposes.

Bromine is used principally in the preparation of bromides, used to a considerable extent in photography and in medicine. It is also used in the preparation of a number of organic drugs and dyestuffs. Bromine is extracted today in large quantities from sea water and apparently the source is inexhaustible.

The chief sources of iodine are brine wells or sea water and the ashes of certain sea weeds. Iodine is used extensively in medicine, especially in the form of tincture of iodine. It also finds an important use in the preparation of iodides and of certain dyes and drugs. The antiseptic iodoform is a compound of iodine with carbon and hydrogen and is similar to the widely used anesthetic—Chloroform—which is a compound of chlorine, hydrogen and carbon. While iodoform and chloroform are valuable drugs, the corresponding compounds of bromine and fluorine are unimportant to medicine. The recovery of iodine from sea water is an important industry.

EXPERIMENT 81—How to Make Chlorine Gas

Equipment: Potassium nitrate, sodium bisulphate, sodium chloride, test tube, test tube holder, alcohol lamp or candle, and measure.

Operation: Put 2 measures of each chemical in a clean dry test tube. Heat the contents gently over a flame for a few minutes. Remove from flame and smell cautiously. The gas which is given off is chlorine.

Result: Sodium bisulphate reacted with sodium chloride to form hydrogen chloride which was oxidized by oxygen from the potassium nitrate to water and chlorine gas.

EXPERIMENT 82—To Show the Bleaching Properties of Chlorine

Equipment: Chlorine gas, blue litmus paper, test tube, alcohol lamp or candle.

Operation: Prepare chlorine gas as in the preceding experiment, placing a small piece of moistened blue litmus paper over the mouth of the tube before heating. Notice on heating that the blue litmus paper turns white, showing that chlorine gas can bleach certain colors.

Result: The chlorine gas reacted with the water on the blue litmus paper, forming hydrochloric acid and oxygen. The free oxygen does the bleaching.

EXPERIMENT 83—How to Make Hydrochloric Acid

Equipment: Ammonium chloride, sodium bisulphate, test tube, blue litmus paper, alcohol lamp or candle, stirring rod, ammonia.

Operation: Put two measures of ammonium chloride and two measures of sodium bisulphate in a test tube. Moisten a piece of blue litmus paper and place over the mouth of the test tube. Heat tube slowly over a flame for a few minutes. Notice that the litmus paper turns red, proving that an acid has been formed. Remove the test tube from the flame and smell cautiously. This is hydrogen chloride gas.

Dip the glass stirring rod in a little household ammonia and hold the rod over the mouth of the test tube. Notice the white fumes of ammonium chloride.

Result: Hydrogen chloride gas when dissolved in water forms hydrochloric or muriatic acid. Commercially, hydrochloric acid is manufactured by heating sodium chloride with sulphuric acid.

HYDROFLUORIC ACID

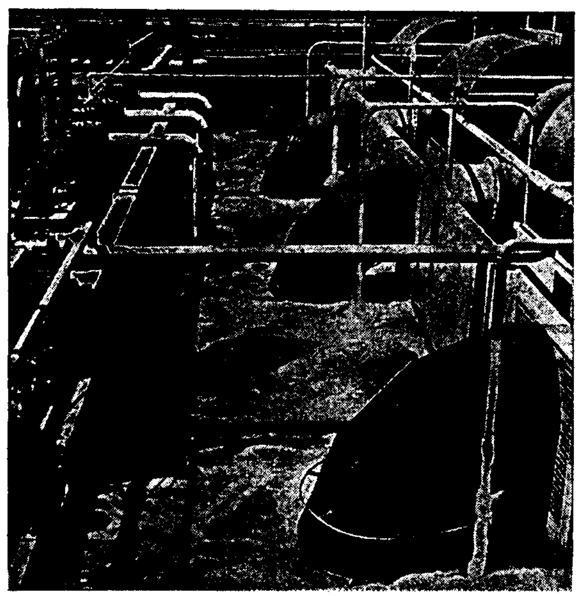
The mineral Fluorspar (CaF_2) is the source of that most interesting halogen acid, hydrofluoric acid (H_2F_2). This acid dissolves glass and thus is used as an etching agent. It must be kept in rubber, wax or lead bottles because it is so highly corrosive. It is important to note here that fluorine is a normal constituent of teeth enamel, and good sound teeth are physiological evidence of proper fluorine nutrition.

BLEACHING POWDER

Bleaching powder or "Chloride of Lime" as it commonly comes on the market, is a compound composed of calcium, oxygen and chlorine. It is prepared by passing chlorine gas over slaked lime, bleaching powder or calcium hypochlorite being formed.

Bleaching powder is a very important compound and has many uses in everyday life. It is a very good bleaching agent and readily gives up its chlorine when treated with an acid. In bleaching paper or rags, the rags are first boiled in an alkali to remove the grease, then placed in a large vat with bleaching powder and sulphuric acid; on removing, the rags are pure white. Cotton cloth is bleached by passing it through alternate vats of bleaching powder and sulphuric acid.

Besides being used as a bleaching agent, chloride of lime is employed as a disinfectant, since it destroys germs. Water is sometimes purified with this compound.



CHEMICALS FROM SEA WATER

Bromine and iodine are extracted from raw sea water with the aid of these four huge centrifugal pumps having a capacity of 230,000 gallons per minute. Acid and chlorine are introduced on the discharge side of these pumps. This plant is located in Texas.

INORGANIC CHEMISTRY

(Continued)

BORON AND THE BORATES

BORON does not occur in nature as a free element. It is found in the form of boric acid in many hot springs, particularly in Italy and California. Boron occurs in large quantities in the desert regions of California and Nevada in the form of its sodium salt, known as borax. It also occurs as the calcium salt. Boric acid is commonly used as a mild antiseptic, as a constituent of talcum powder, and as a preservative.

EXPERIMENT 84—Test for Boric Acid

Equipment: Boric acid, alcohol, test tube, saucer.

Operation: Dissolve 2 measures of boric acid in a test tube ⅓ full of alcohol. Now pour this solution into a saucer, darken the room and light the alcohol. Notice that the alcohol burns with a green flame.

Result: This green flame is characteristic and is a test for boric acid.

EXPERIMENT 85—Examination of Talcum Powder

Repeat Experiment 173 using instead of the boric acid 3 or 4 measures of a talcum powder. Some talcum powders contain boric acid. Test three different commercial grades of talcum powder for boric acid.

BORAX GLASS

When heated, borax swells up to a bulky mass, loses its water of crystallization and then melts to a clear glass. This glass readily dissolves various metallic oxides which impart characteristic colors to the glass. This property is used in testing for certain metals. For this same reason borax is used as a flux in brazing or hard soldering to remove the oxides from the surface of the metals to be joined.

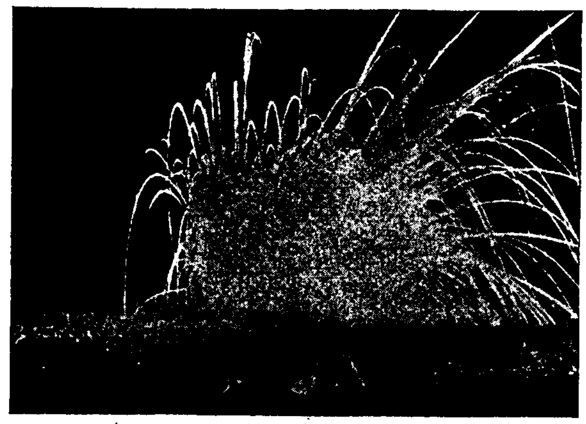
PHOSPHOROUS AND THE PHOSPHATES

Phosphorus never occurs free in nature but is found in combination with oxygen and metals as derivatives of phosphoric acid. It occurs most extensively as calcium phosphate. All fertile soils contain calcium phosphate. Since it is essential to plant growth, it is an important constituent of fertilizers, the soluble calcium monophosphate being used. The bone of animals is largely calcium phosphate.

Phosphorus exists in two forms. Yellow phosphorus is a transparent, waxlike solid which often takes fire in air at ordinary temperature. It is always kept under water. It is used extensively in modern warfare and is very poisonous.

Red phosphorus is a chocolate red amorphous powder and is quite stable in air at ordinary temperature.

The phosphates of most metals can be precipitated from solution of the salts since most phosphates are insoluble in water. The phosphates of the alkali metals are soluble in water.



White Phosphorus grenade lands in gun emplacement

TREATMENT OF PHOSPHOROUS BURNS

A phosphorus burn should first be washed with a large volume of water and kept wet. Unless all small particles of unburned phosphorus can be removed with the forceps, it may enter the blood and cause phosphorus poisoning. Within two minutes after the accident the burn should be washed alternately with 1% solution of copper sulphate and a solution of sodium bicarbonate, avoiding at first the use of oily liniments and salves. The wound can then be treated in the usual manner with salves and liniment.

Free yellow phosphorus previously dipped in 2% copper sulphate solution can be introduced into a wound sac under the skin of a guinea pig without injuring the health of the animal.

THE ALKALI METALS

The alkali metals are lithium, sodium, potassium and rubidium and they are all very similar in their chemical behavior. The two most common are sodium and potassium. These metals are called alkali metals because they are a part of the compounds which are known as alkalies or bases, such as caustic soda (sodium hydroxide) and caustic potash (potassium hydroxide).

Because of their extreme reactivity none of these metals occur free in nature. They are found as compounds, distributed in sea and mineral water, salt beds and rocks. The free metals are soft, with a silvery luster, but tarnish quickly in air because of their reaction with oxygen. They are kept in kerosene out of contact with the air. They react vigorously with water, setting free hydrogen from the water forming alkaline solutions.

Some of the more important compounds are sodium hydroxide, the chief constituent of ordinary lye; sodium chloride, or common salt; sodium carbonate, or washing soda; sodium bicarbonate, or baking soda, potassium nitrate known in trade as salt peter, sodium nitrate and sodium and potassium sulphates.

POTASH

Potassium carbonate (K_2CO_3) commonly known as *potash*, is a very important and essential compound of the element potassium. It is essential to plant growth, and is a component of all fertilizers. It is absorbed by growing plants from the soluble inorganic compounds present in the soil and in commercial fertilizers. An excellent source of potash is wood ashes and potash is obtained easily by leeching out with water the soluble material in wood ashes.

ALKALINE EARTH METALS, CALCIUM, STRONTIUM, BARIUM

The term "alkaline earth" was originally applied to the oxides of these metals because they resemble both the alkalis and the earths, the latter term being applied to oxides of aluminum and iron.

These metals do not occur free in nature, but largely as carbonates and sulphates. They are light and active, resembling each other closely in physical and chemical properties. They react with the oxygen in the air and decompose water, liberating hydrogen, similar to the alkali metals.

Calcium carbonate is found in nature in large quantities as limestone or marble. Also gypsum (calcium sulphate) and phosphate rock (calcium phosphate) occur extensively.

Lime or calcium oxide is prepared from limestone by heating in large furnaces or kilns. It is used in making calcium hydroxide and slaked lime. (See Exp. 87).

EXPERIMENT 86—Burning Limestone to Quick Lime

Equipment: Limestone.

Operation: Crush small pieces of limestone with a hammer. Place some of this material—2 teaspoonsful—on a small tin can lid and heat over a hot gas flame. Heat until the limestone becomes white hot and after twenty minutes' heating allow to cool. The compound remaining is quick lime or calcium oxide.

Result: Limestone or calcium carbonate, upon heating, decomposes into calcium oxide and carbon dioxide.

EXPERIMENT 87—Slaked Lime or Calcium Hydroxide

Operation: To a test tube add ½ inch of powdered lime and 5 or 6 drops of water. Notice that the lime puffs up after a few minutes and appears to be perfectly dry. This is known as slaked lime or calcium hydroxide.

Now fill the test tube ⅓ full of water and shake the contents of the tube. Filter off the liquid and test it with red litmus paper.

Result: This liquid solution of calcium hydroxide shows a weak base reaction (Test). It is sold by the druggist as *Lime Water.*

COMMON MASON'S MORTAR

Mortar, which is used in stone foundations for buildings, is made by mixing together water, sand and slaked lime. On exposure to the air mortar sets or becomes hard. This is because the carbon dioxide in the air reacts with the slaked lime to form an insoluble calcium carbonate or limestone.

EXPERIMENT 88—Making Mortar

Take three measures of calcium oxide and mix thoroughly with three measures of sand. Then add a few drops of water to make a paste.

Spread this paste on a board and allow it to stand for several days. Notice that it soon becomes hard.

SULPHIDE LUMINESCENCE

Calcium sulphide when it contains traces of sulphides of some of the other metals has the property of absorbing light from a luminous body and then giving up the light in the dark. It is prepared by heating calcium sulphate with carbon.

EXPERIMENT 89—Pharaoh's Serpents

Equipment: Potassium dichromate (purchased separately), potassium nitrate, sugar.

Operation: Mix two parts potassium dichromate, (this can be secured in a drug store), one part of potassium nitrate and one part of sugar. The ingredients must be powdered separately and then mixed. Small paper cones can be made and filled with the compound which has been moistened either with alcohol or water. When dry, light at the top. As the cone burns "snakes" will form.

EXPERIMENT 90—Red Flame from Sawdust

Boil some sawdust or woodshavings in a cup of water containing a teaspoonful of potassium nitrate. When it is dry it will burn with a white-yellow flame, sizzling as it burns.

ALUMINUM, ZINC, MAGNESIUM

These are three very important metals and find wide application in industry.

Aluminum is valuable because it is very malleable, is easily cast, is tenacious, and more rigid than the same weight of other metals. Its uses are many. It is used for common household aluminum ware. It is used extensively in the manufacture of airplanes on account of its lightness and durability. Recently it has been used to construct the gondolas of the stratosphere balloons. Aluminum is used as a conductor in electric lines.

Aluminum forms valuable alloys with steel or magnesium.

The compounds of aluminum are important. The silicates are used extensively in manufacture of cement, brick, tile, earthernware, pottery, and porcelain ware. Several valuable gems contain aluminum. The ruby, sapphire, and topaz are transparent crystals of aluminum oxide, containing small amounts of certain metallic oxides which impart the color.

The alums, which are double salts containing aluminum sulphate, are used as mordants in dyeing and printing. Because of its acid reaction in water, alum is used in some baking powders.

MAGNESIUM

Magnesium occurs abundantly in nature as magnesium carbonate. Metallic magnesium is very important commercially because of its extreme lightness and strength. It is now a rival of aluminum in construction demanding a light metal. Many of the large trucks used for transporting new automobiles are constructed of a magnesium alloy, "Dow metal" which is being extensively used. The airplane industry has demanded an extensive use of magnesium metal and its alloys. There is a great abundance of magnesium in nature furnishing an unlimited source of this important element. Its uses are daily being expanded and is developing into an enormous industry.

When ignited in air magnesium burns with a brilliant white light. For this reason the powdered metal, mixed with potassium chlorate (or potassium nitrate), is used as a flashlight powder in photography, as well as in fireworks and signal flares.

The well-known milk of magnesia is magnesium hydroxide; the drug epsom salts is magnesium sulphate. They are both commonly used household remedies.

SILICATES

EXPERIMENT 91—Nickel Silicate

Equipment: Nickel ammonium sulphate, water glass, test tube, candle or alcohol lamp.

Operation: Place two measures of nickel ammonium sulphate in a test tube and fill the tube half full of water. Warm the solution for a few moments to completely dissolve the solid and immerse the tube in cold water to cool it again.

Add two or three drops of water glass and you will get a beautiful green precipitate of nickel silicate.

Result: This is characteristic of nickel silicate.

EXPERIMENT 92—Silicic Acid

Equipment: Sodium bisulphate, sodium silicate, water glass, test tubes, measure.

Operation: Put ½ inch of water glass in a test tube and add water until the tube is one-quarter full. Shake to mix the liquids.

In another test tube put four measures of sodium bisulphate and fill the tube one-third full of water. Shake until the solid is completely dissolved.

Pour the sodium bisulphate solution into the water glass. A jelly-like precipitate will form, and in a few minutes all the liquid in the tube will become solid.

Result: This jelly-like precipitate is Silicic acid. It is a typical collodal substance having no crystalline form.

SODIUM SILICATE (WATER GLASS)

EXPERIMENT 93—Sodium Silicate Paste

Paint a thin film of water glass on a sheet of paper and let it dry for 15 or 20 minutes. Note the smooth transparent glass-like film which results.

Paste together two sheets of paper or two blocks of wood, using water glass as the adhesive. You will find that it makes an exceptionally strong paste and it is often used for this purpose.

Part IX $HC_2H_3O_2$

ELECTRO - CHEMISTRY

BEFORE we discuss the part electricity takes in chemistry we must first know a little something about electricity. Electricity, like heat, is a form of energy. About 100 years ago very little was known about the role that electricity played in chemical reactions. Today matters are quite different. Electricity and chemistry are very closely related. Many large and important industrial concerns are engaged in manufacturing materials involving the use of electro-chemical reactions.

Today chlorine gas and caustic soda are manufactured by passing an electric current through salt water. From the chlorine gas we obtain bleaching powder. Metals are extracted from their ores by passing a current through their molten or fused salts. Nickel-plating, copper-plating and gold-plating are done by passing a current through a solution containing salts of these metals. The success of these important industries and many others is based on the fact that electricity possesses the power of decomposing chemical compounds.

On the other hand, we can show the relationship between chemistry and electricity in another way. We have already said that electricity is a form of energy. Now, in most chemical reactions, heat is liberated as the form of energy. However, under proper conditions, the energy of certain chemical reactions is liberated in the form of electricity. For example, if we put a copper plate and zinc plate in a solution containing an acid and connect the two plates with copper wires we find that a current of electricity is produced. A reaction takes place in which electricity is the form of energy liberated. Use is made of this fact in the manufacture of the different types of electric cells and batteries. Batteries are simply cells connected together in series in order to produce a stronger current. There are several types of cells, all of which come under two classes, namely, the primary cells, which include both the dry and wet cells, and the secondary cell or storage battery, as it is called.

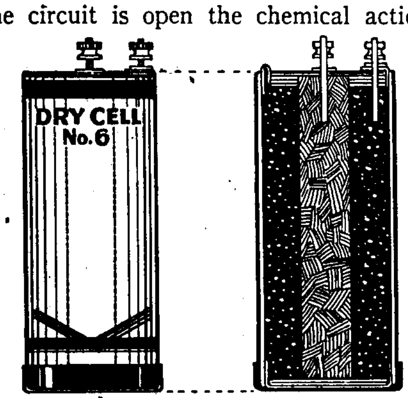
Finally, we will mention a third relationship of electricity to chemistry. That is, the part electricity plays in furnishing heat to produce chemical change. A good example of this is the manufacture of graphite from carbon by means of the electric furnace. Also the manufacture of calcium carbide for the production of fertilizers, of nitric acid from nitrogen in the air, and many other important industries depend upon the heat generated by the electric current for their success.

Before we discuss any of the different types of cells we will first consider the parts that go to make up a primary cell. These are, first, the jar which holds the solution and the elements; second, the solution of electrolyte, as it is commonly called; third, the cathode or negative electrode, which is usually made of the element zinc; and, fourth, the anode, or positive electrode, which is usually the element carbon.

THE DRY CELL AND HOW IT IS MADE

The dry cell is a very common type of cell and millions of them are used for bell, telephone and other purposes. The jar of the dry cell consists of a cup of sheet zinc which serves as the negative electrode (below). A binding post is fixed at the top of this cup. The electrolyte in the dry cell consists of an active paste which consists usually of one part of zinc chloride, one part of zinc oxide, one part of ammonium chloride or sal ammoniac, three parts of plaster of Paris, two parts of manganese dioxide and one part water, all by weight.

The cell is prepared as follows: the zinc cup is filled to within one-half inch of the top and a carbon rod containing a binding post on the upper end is pushed down into the paste to within an inch of the bottom. Melted pitch is then poured over the paste until it is even with the top of the cup. The pitch is then allowed to cool, and the cell is ready for use. It is only when you close the circuit, that is, connect the two binding posts with a wire, that you start chemical action within the cell. When the circuit is open the chemical action stops.



HOW THE DRY CELL WORKS

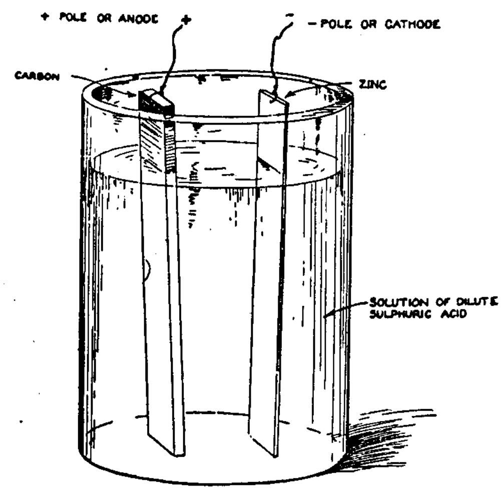
After the cell is made and copper wire is attached from the carbon post to the zinc post a current of electricity passes through the wire. This is due to action of the active paste upon the zinc electrode. When the circuit is opened, that is, when the carbon and zinc posts are disconnected, the action stops.

Now the chemical action which takes place within the cell is as follows: The ammonium chloride and water react to form hydrochloric acid which attacks the zinc. The zinc goes into solution in the form of positive ions, which are atoms or groups charged with electricity. When this happens the zinc electrode assumes an electro negative condition. The positive zinc ions unite with the negative chlorine ions of the hydrochloric acid to form zinc chloride with the formation of positive hydrogen ions. Now, the positive hydrogen ions move to the carbon electrode where they lose their charge and become gaseous hydrogen. Therefore, when the circuit is closed the chemical action, which takes place, keeps the zinc pole or cathode negatively charged and the carbon pole or anode positively charged. The flow of electricity is always from the negative zinc pole to the positive carbon pole through the solution, and from the positive carbon pole to the negative pole through the wire.

You might ask the question, what happens to the hydrogen gas when chemical action takes place within the cell? The hydrogen gas as fast as it is formed at the carbon electrode is oxidized to water by oxygen from the manganese dioxide. This brings up the phenomenon known as polarization. By polarization is meant the cutting down of an electric current, due to the lowering of potential between the carbon and zinc poles. Polarization in a cell is caused by formation of bubbles of hydrogen gas clinging to the carbon electrode, thereby producing less surface. To prevent this, manganese dioxide is used, which, as already stated, oxidizes the hydrogen gas to water.

THE WET CELL

The wet cell or other type of primary cell is very similar to the dry cell. The electrolyte instead of being a paste, as in the dry cell, is a solution. The positive and negative electrodes are of carbon and zinc or of copper



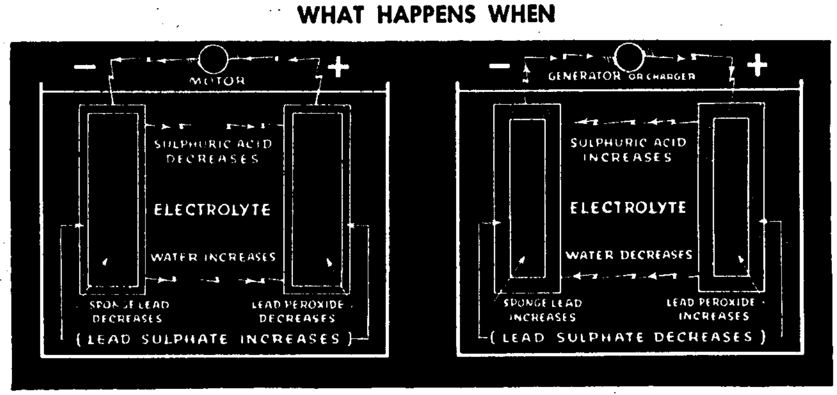
and zinc (above). The principles involved in the formation of a current due to chemical action in the wet cell is the same as that in the dry cell, although there are several types of wet cells.

THE STORAGE BATTERY

Another type of electric cell or storage battery is a little more complicated than the dry or wet cell. The storage battery consists of a number of secondary cells. It is used largely for running electric power plants, electric industrial trucks, submarines, telephone and telegraph work, automobile starting and ignition, etc.

The storage battery does not generate a current of electricity like a primary cell by a direct chemical action. It is charged by a current of electricity, after which it will deliver a current until the cell is run down.

WHAT HAPPENS WHEN



BATTERY DISCHARGES

A BATTERY IS RECHARGED

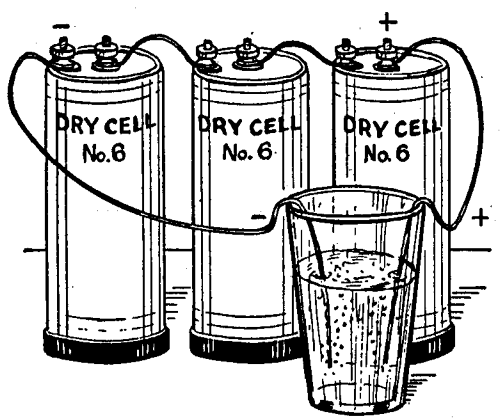
Courtesy Willard Storage Battery Co.

EXPERIMENT 94—How to Test a Battery of Dry Cells

If you have two or three dry cells around the house and wish to test their strength, perform the following experiment with them:

Connect the cells together in series by attaching small pieces of copper wire to the carbon binding post of one cell and zinc binding post of the next cell. Now connect a longer piece of wire to the free carbon post and another long piece to the free zinc post. Before connecting, clean the ends of all the copper wires by scraping with a knife blade.

Now dissolve two teaspoonfuls of sodium chloride (table salt) in a tumbler two-thirds full of water and put the ends of the copper wires, which should be clean, in this solution. Notice whether bubbles of gas appear on the ends of the wires in the solution. By setting up the same number of new cells in the same manner and comparing the amount of gas produced with that produced from the old cells you can tell whether the old



cells (above) are of sufficient strength to be of value in performing the following experiments.

If no gas is formed from the old cells when the preceding experiment is accurately performed, the cells are worn out and you will have to obtain some new cells. Two or three dry cells connected in series will give you sufficient current for performing most of the experiments as outlined.

EXPERIMENT 95—How to Show the Direction of a Current

Equipment: Measure, nickel ammonium sulphate, two or three dry cells.

Operation: Dissolve 8 measures of nickel ammonium sulphate in a tumbler one-third full of water.

Now connect two or three dry cells in series as shown in Experiment 94 and put the ends of the wires leading from the positive and negative poles of the battery into the nickel solution. Notice that very soon the wire attached to the negative pole or zinc electrode is coated with metallic nickel or is being nickel-plated.

Now disconnect the two wires from the battery and attach the wire which was plated with nickel to the positive carbon pole and the other wire to the negative zinc pole. Put the ends of the wires again in the solution. Notice that the nickel is soon dissolved from the plated wire and is deposited on the other wire which is attached to the negative pole.

Result: This proves that the current always flows from the positive pole to the negative pole and that the metal deposited always follows the direction of the current. This experiment will determine the positive or negative wire, as only the negative wire will show plating.

ELECTROPLATING

The very simple process of transferring metal from one object to another by chemical and electrical means is called electroplating. All of the silverware in use is plated by the same process you are going to use. By this method objects are copper-plated, silver-plated and gold-plated.

Besides its use in plating, the process is used in the purification of certain metals. Copper, for example, is separated from its impurities in this manner. At a recent chemical exhibit in New York, a slab of copper five feet square and six inches thick was placed on exhibition which had been purified not by removing impurities from the copper, but by removing the copper from the impurities.

In electroplating, the electrolyte or bath always consists of a solution of the salt of the metal to be deposited or plated on the object. Now, as to the action which takes place when an object is electroplated, let us first consider the nature of solutions of metallic salts when a current is passed through the solution. Salts are made up of metallic elements and non-metallic elements or groups. When in solution the metallic elements become ions and have positive electric charges. The non-metallic elements or groups on the other hand also become ions and have negative charges. Now suppose we pass a current through a solution of copper sulphate. The metallic copper ions which are positively charged are attracted to the negative pole to which is attached the object to be plated. On reaching the object to be plated the copper ions lose their charge, become atomic of metallic copper and as such are deposited in a smooth thin layer upon the object.

The non-metallic sulphate groups which are negatively charged are attracted to the positive pole to which, in the case of copper-plating, is attached a sheet or bar of copper. Upon reaching the positive copper pole the sulphate groups lose their charge, become molecular sulphate, having the properties of a strong acid grouping and dissolve the copper to form copper sulphate which goes into the solution. The amount of copper which goes into solution in this way is exactly equal to the amount of copper which is deposited upon the object to be plated. You can see, therefore, that the concentration of the copper sulphate solution is always the same as long as there is any copper left at the positive pole.

The preceding action may be expressed a little more clearly in the form of an equation, thus:

Salt + electricity = metal (of salt) goes to the object to be plated (cathode —)
non-metal (of salt)— goes to the metallic plate (annode +)

By using different kinds of salts and plates of different metals we can plate with almost any metal although some metals plate easier than others.

EXPERIMENT 96—How to Nickel-plate

Equipment: Sodium carbonate, nickel ammonium sulphate.

Operation: The object to be nickel-plated must be free of oil, grease and varnish. This can be done by boiling it in vinegar or a solution of sodium carbonate for several minutes. Then wash in clean water.

Dissolve one spoonful of nickel ammonium sulphate in a tumbler half full of water. Now attach the iron, copper or brass object to be nickel-plated to the negative wire and an iron nail to the positive wire. Immerse these in the solution.

Result: Notice that soon the object attached to the negative wire which goes to the zinc post is covered with a coating of nickel.

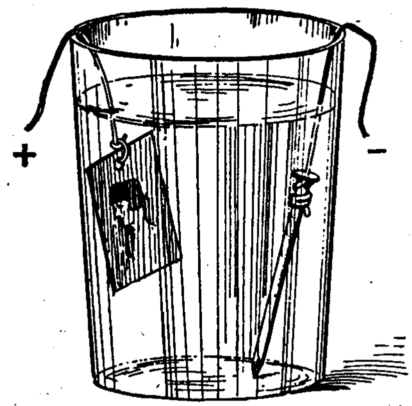
EXPERIMENT 97—How to Etch on Steel

Equipment: Paraffin, nickel ammonium sulphate, tumbler.

Operation: Steel or iron can be etched the same as copper. Procure a piece of sheet steel or iron and after coating it with paraffin trace the design upon it. Note: A knife blade or any small steel object may be etched in this way. Then connect it to the *positive* wire leading to the carbon post and attach a bright nail to the negative wire leading to the zinc post of the battery.

Now place the steel and the iron nail in a solution of nickel ammonium sulphate made by dissolving one spoonful of the compound in a tumbler half full of water.

Result: Allow the current to pass through the solution for several hours and then remove the steel and melt off the paraffin. Notice that the design is etched upon the steel.



EXPERIMENT 98—Tin-Plating by Contact

Equipment: Tartaric acid, measure, alcohol lamp or candle.

Operation: Dissolve 6 or 8 measures of tartaric acid in a tin cup half full of water. Now place in this solution a penny which has been cleaned by boiling for several moments in a little vinegar, or sodium carbonate solution made by dissolving 4 measures of sodium carbonate in a test tube ½ full of water.

Result: Put the tin cup on the stove and allow the water to boil off. Notice that after several minutes the penny will gradually become coated with a bright silvery plating of tin.

EXPERIMENT 99—How to Clean Silverware Electrolytically

Equipment: Table salt, old aluminum pan, silverware to be cleaned.

Operation: If you have any silverware which is stained dark by exposure to the air you can easily remove the stain, which is silver sulphide, by treating the silverware as follows:

Obtain an old aluminum pan and place the silver to be cleaned in the pan. Now cover the silver with a solution of common salt or baking soda made by dissolving two spoonsful of the salt in each quart of water used. Now place the pan on the stove and allow the solution to boil for two minutes. Remove the silverware and wash it with fresh water. Notice that the black stains are removed and the silver is bright and clean.

Result: The black stain or silver sulphide was reduced by the chemical action taking place in the solution. A feeble electric current was formed in which the aluminum pan acted as the negative pole and the silverware as the positive pole. The electrolyte in this case was the solution of common salt or baking soda.

The metal silver cleaners which you probably have seen advertised on the market are simply metals of aluminum or zinc. The process of cleaning silverware with these cleaners is the same as that used in this experiment.

Part X

ORGANIC CHEMISTRY

CARBON

THE science of Organic Chemistry is based on our knowledge of the properties and reactions of the compounds of the element—*carbon*. No other element in nature is capable of uniting with other elements to form so many compounds of theoretical and practical interest. It may be called the element of life; all of the manifestations of life are dependent on the existence of some compound in nature containing this element.

The element *carbon* is found in nature in free forms in different allotropic modifications. In other words, "variations in physical properties shown by many chemical elements without change or alteration of chemical composition, and this is characteristic of the element carbon."

FORMS OF CARBON

The diamond is practically pure carbon, while ordinary coal and graphite contain different percentages of other substances besides carbon including much mineral matter. Naturally occurring compounds of carbon are of very wide occurrence in nature and are found in the form of gases, liquids, oils and solids. Carbon dioxide is the most familiar, gaseous combination of carbon. Manufactured illuminating gas, natural gases from wells and petroleum in all the modifications met with in nature are all composed chiefly of organic compounds of carbon and hydrogen.

Of the inorganic minerals containing carbon, the carbonates, especially calcium carbonate, constitute a very large proportion of the natural rocks and some form of mineral carbonate is found in most localities. The building stone, marble, is a very pure form of calcium carbonate. Carbon constitutes a large percentage of all living organisms, both animal and plant, and is represented in such organic products as proteins, fats, sugars and natural oils. All such products are widely used by man as food and for the manufacture of useful commercial products. At the present time more than 450,000 organic compounds are known, and the possibilities of new creations as the science of organic chemistry is developed are unlimited. Our present world could not exist without this element.

Coke is an artificial or a modified form of impure carbon, and is used as a domestic fuel, for operating steam boilers, and also in smelting processes for refining ores. It is made by heating bituminous or soft coal at a high temperature until all volatile products in the coal have been expelled. This heating process is conducted without access of air, and is carried out on a commercial scale in large ovens. The volatile or gaseous products are refined and constitute the raw materials for the manufacture of illuminating gas and low boiling hydrocarbons like benzine and toluene.

Every boy and girl is familiar with ordinary charcoal. This is a form of carbon produced by the destructive distillation of organic substances such as wood and sugar, and even bones of animals. Destructive distillation is heating without access to air as in the manufacture of illuminating gas. The quality of wood charcoal obtained is dependent on the kind of wood used. It is known, for example, that the destructive distillation of cocoanut shells produces a very efficient form of absorbent carbon meeting the exacting requirements of the gas mask. This form of carbon is referred to in the trade as "activated carbon or charcoal." If air was admitted during the distillation process, the charcoal and gaseous products would be burned up completely with formation of carbon dioxide gas and water.

Another form of carbon is *graphite,* the black substance which forms the core of lead pencils. It is sometimes referred to as lead, but this is not correct. Graphite is a modified form of carbon, made by subjecting carbon to a very intense heat. At about 4000°C. carbon vaporizes, and this vapor on condensing forms graphite. Graphite is used in the manufacture of crucibles, as a lubricant, as a protective covering for metals such as stove polish, and in the manufacture of lead pencils.

COLLOIDAL GRAPHITE AS A LUBRICANT

The high temperatures encountered in many present-day industrial processes have emphasized the need for special lubricants and *graphite* has given satisfaction in such applications. The conventional lubricants designed for use under conditions of high temperatures have usually been heavy greases or high viscosity oils. Such products, however are unable to withstand the high temperatures which often exist in many industrial operations. Oils, when subjected to high heat, decompose or distill off, and greases are consumed or carbonize leaving behind a non-lubricating residue.

Obviously, then, a lubricant which will perform under such conditions would possess unusual properties. Electric furnace graphite, colloidally dispersed in a suitable carrier fluid, adequately meets these conditions as a lubricant. For example, it finds use in operating baking and enamelling oven mechanisms, conveyor chains and wheels operating in ovens at 1300°F for hours at a time, in bottle machines, in incandescent lamp machines, and as die casting and forging lubricants. Graphite diffused in castor oil is highly desirable as a mold lubricant on automatic type-setting machines. *Searchlights:* A special mixture of colloidal graphite in water is a standard lubricant for the carbon feed gears and bearings which automatically maintain the arc in high intensity searchlights. Many other applications might be mentioned.

LAMP BLACK

Lamp black, an amorphous form of carbon, finds wide application in the trades. It is used in rubber as a toughening agent; in printers' ink as a pigment; in paints, stove polishes and lacquers as a pigment; on typewriter ribbons and carbon papers as a black coloring matter. Bone black, an amorphous carbon produced by destructive distillation of bones, is used as a deodorizing and decolorizing agent.

THE MANUFACTURE OF ILLUMINATING GAS

Cylindrical ovens of fire clay (B — see below) are filled automatically with soft coal. These ovens are then closed tightly to prevent entrance of air. Under these ovens is a hot fire (A). The heat decomposes the coal into gases, liquids and coke. The gases contain impurities such as hydrogen sulphide, carbon dioxide, ammonia, tar, benzol, toluol and water. These impurities are removed before the gases are passed into the tanks.

The first step towards purification of the gases is as follows: The gases pass from the oven through a pipe, and bubble through running cold water contained in the lower half of a large pipe (C). Here coke dust (carbon), tar oils and ammonia are removed.

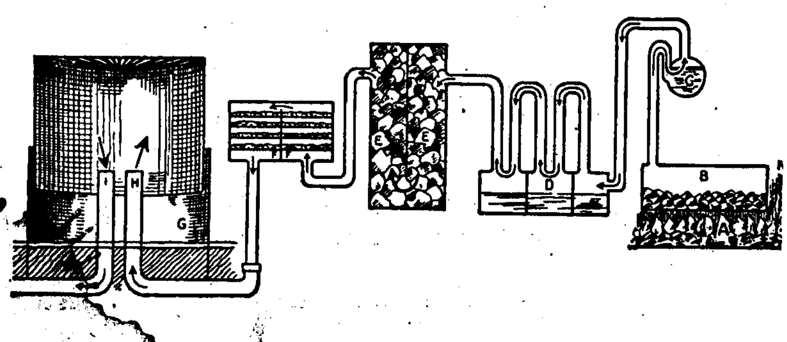
Second, the gases are passed through an arrangement (D) which consists of several hundred feet of pipe. This acts as a condenser and cools the gas down to ordinary temperature and condenses the liquid products of the distillation.

Third, the gases pass from the condenser through a "scrubber" where they are washed and cleaned. The scrubber (EE) is a large iron tank filled with coke, crushed rock, wood and scraps of tin, the object being to expose a large surface to the gas. A spray of water is introduced at the top of the scrubber and the material filling it is thus kept moist. The remainder of the tar and ammonia salts are here removed.

Fourth, the gases are passed on to the "purifier" (FF), a rectangular box filled with layers of quick lime, which absorbs water, carbon dioxide and hydrogen sulphide.

After the moisture is removed, the gases, which now consist of hydrogen, nitrogen, marsh gas, ethylene gas, acetylene and carbon monoxide are delivered immediately into the large gas tanks (G). These tanks are constructed in telescopic fashion so that the quantity of gas regulates and controls the size of the tank. From the tank it is pumped through gas mains to the homes of the consumers. (H) is the entrance pipe. (I) is the exit pipe. (K) is the flue or chimney for the fire.

Every product formed by the heat decomposition of coal is saved in the gas plant. All the hydrocarbon by-products are separated and condensed and appear as benzene and toluene which are very valuable.



EXPERIMENT 100—Preparation of Charcoal—Pyro-ligneous Acid

Equipment: Test tube holder, blue litmus paper, alcohol lamp or candle, test tube.

Operation: Break up a few toothpicks or pieces of wood and place them in the bottom of a test tube. Now put a piece of moistened blue litmus paper over the mouth of the test tube and heat over a flame. Notice that the paper turns red, proving that an acid is evolved. This acid, called pyro-ligneous acid, is essentially acetic acid the same acid which is also present in vinegar. Here it is formed in the process of fermentation regulated by the action of bacteria on sugar or sugar contained in fruit.

Now insert a perforated cork with delivery tube, continue heating and light the gas that comes off. *Notice that it will burn.* This gas is similar to that obtained from the distillation of coal.

When no more gas is evolved, allow the tube to cool, then empty the contents of the test tube on paper. This is charcoal and is practically pure carbon.

When green pine (or green cedar wood) is distilled, turpentine and tar oils are derived from it. The turpentine is the volatile oil possessing a pleasant odor which passes off as a vapor. Tar oils are heavy resinous oils, brown in color, such as you may see at the bottom of your test tube. These are phenolic in nature.

Result: To summarize, when coal is heated in the absence of air, coal gas, ammonia and coal tar are obtained. The gas is used for lighting and heating. The ammonia which is derived is purified and finds many uses. From the coal tar are derived intermediates from which aniline dyes, disinfectants such as carbolic acid, high explosives and many other valuable products are manufactured.

It can be seen, therefore, that enormous industries are based upon this process of heating wood and especially coal without admission of air.

THE SMOKING OF HAMS AND MEATS

Pyro-ligneous acid is contained in the smoke of smoke houses where hams and other foods are cured. In the pioneer days of New England agriculture, a smoke house was a common feature of the farmer's equipment. Usually this was built out-of-doors, but in many cases installed in the attic of the farm house. It is due to the aseptic action of this organic acid in the smoke that pork and beef products are preserved. The probable action of the gradual deposition of pyro-ligneous acid upon a ham is to gradually kill all of the bacteria or germs which are the cause of decay. The acid also imparts that distinctive taste which is characteristic of smoked foods.

Meats can be cured over night, while in the smoke-house several days are required. This quick curing is a rather new method and is not practiced except experimentally.

The ham is painted with pyro-ligneous acid, which seeps into the meat. It has been declared that the quick cured ham is as edible as the slow cured ham and keeps just as well.

EXPERIMENT 101—Decolorizing Vinegar

Operation: Fill a test tube one-quarter full of vinegar which has a brownish or yellow color. Add two measures of charcoal and shake with the solution for four or five minutes.

Result: Now separate the charcoal by filtering, and you will find that the color of the vinegar is lighter. The vinegar can be made almost colorless by repeating this several times.

EXPERIMENT 102—Absorbent Charcoal from Ground Cocoanut Shells

Grind up a piece of dry cocoanut shell and thoroughly bake the material in a copper oven. This should be thoroughly carbonized by this treatment. After baking, grind the particles of dried shell to a powder by rubbing in a mortar. Test the efficiency of this powder as a decolorizing and absorbing agent.

EXPERIMENT 103—Absorbent Charcoal from Butternut Shells

Crack some butternuts and save the meat of the nuts. Then take the shells and thoroughly crush them and finally carbonize by heating in a copper oven. After this baking, then grind to a powder in a mortar. Test the efficiency of this powder as a decolorizing and absorbing agent.

EXPERIMENT 104—Preparation of Absorbent Charcoal from Hickory Nut Shells

Follow same directions as given for Experiment 103, using the shells of hickory nuts.

EXPERIMENT 105—Preparation of Absorbent Charcoal from White Birch Wood

Carbonize some small pieces of white birch wood in a copper oven. After baking, then grind to a powder in a mortar. Test the efficiency of this powder as a decolorizing and absorbing agent.

EXPERIMENT 106—Surface Tension and the Rubber Band

Operation: Float a thin rubber band on a dish of water and touch the water inside the band with a wire or tooth pick which has been dipped in oil. The band will snap out, forming a circle. Now apply oil to the water outside of the band and the band will again resume its original shape.

Result: This experiment illustrates the effect of surface tension which tends to make liquids assume those forms which expose the least surface for a given volume.

EXPERIMENT 107—Changing the Specific Gravity of Charcoal

Equipment: Small piece of wood charcoal, test tube, test tube holder, alcohol lamp or candle.

Operation: Wood charcoal floats on water. Tie a weight on a piece of charcoal with a thread so that it will sink and place it in a test tube one-quarter full of water.

Result: Boil the water for several minutes, then remove the weight from the charcoal and you will find that it no longer floats.

Explanation: This is due to the fact that the air is driven away from the pores of the charcoal by boiling and water takes its place. This experiment illustrates why wood becomes waterlogged and does not float.

CARBON DIOXIDE OR CARBONIC ACID GAS

When carbon or any combustible compound of carbon is burned, the carbon is converted into carbon dioxide as the final product of oxidation. This gas is heavier than air, and is the most commonly known of all carbon derivatives. While it does not support combustion, it does serve a valuable purpose in both human and plant economy. In the plant kingdom, carbon dioxide is absorbed from the air through the cells of the leaves and furnishes the source of carbon for building up plant tissues. The transformation of carbon dioxide in the plant is brought about under the agency of the sun's rays and the influence of the green chlorophyll of the leaves of the plant.

Function of Carbon Dioxide in Breathing

While carbon monoxide is poisonous, carbon dioxide is a harmless gas. It is a waste product thrown off by the lungs during respiration. It has been proved that carbon dioxide in the lungs is responsible for stimulating the respiratory center in the process of breathing. The deep and rapid breathing from extensive violent physical exercise, like baseball and football, is not due directly to the need of oxygen, but rather to the need of eliminating carbon dioxide from the lungs. The increase of the carbon dioxide production is a measure of the work being done under violent exercise. Use is made of carbon dioxide for administration to patients suffering from hiccoughs. Also to increase the breathing rate after an anesthesia and even after carbon monoxide poisoning. While carbon dioxide is not ordinarily considered poisonous, it can, however, be responsible for death, because it will not support combustion. The ordinary procedure of testing the air in a mine or a deep well, or any building or enclosed place having poor ventilation, with a lighted candle has proved to be very wise in many cases.

EXPERIMENT 108—Weight of Carbon Dioxide Gas

In collecting carbon dioxide, we have already made use of the fact that carbon dioxide is much heavier, that is, much denser than air. Its high density makes it possible to pour it from one vessel to another almost like a liquid.

Equipment: Baking soda, candle, vinegar, milk bottle, tumbler.

Operation: Obtain a quart milk bottle and in it place two teaspoonsful of baking soda. Now set a short candle upright in the bottom of a tumbler or jelly glass, using a little melted wax to make it stick. The wick should come not higher than half way up the tumbler. Pour about half a glass of vinegar into the milk bottle, and while the carbon dioxide gas is being formed and pushing the air out of the bottle, light the candle. Now hold the milk bottle over the glass and pour carbon dioxide out of it on the candle, just as if you were pouring a liquid, but be careful not to let any of the vinegar come out.

Result: When you have filled the glass up to the level of the flame with carbon dioxide, the candle will go out, for the oxygen supply will be cut off by the heavy CO_2 gas. This gas will not support combustion.

Carbon dioxide, because of its relatively high density, tends to collect in abandoned mines and old wells, therefore such places should always be tested by lowering a lighted candle into them before anyone descends. Sometimes carbon dioxide formed by fermentation accumulates in silos; suffocation may result from entering.

GROWTH OF PLANTS AND UTILIZATION OF CARBON DIOXIDE

The leaves of green plants have the power of combining carbon dioxide and water to form sugars, starches and cellulose. This reaction cannot take place unless energy in the form of light is supplied to the leaf, and hence the process by which the plants use carbon dioxide and water is called "photosynthesis". Without this reaction animal life as we know it would be impossible. It was by means of photosynthesis that the energy of the sun's rays millions of years ago was stored in the form of coal. The carbon in coal was originally obtained from the carbon dioxide of the air by the energy of sunlight shining on green leaves.

Thus plants convert the carbon dioxide of the air into substances which can be used for food or fuel, and then the food is oxidized in the bodies of animals and the fuel is burned in a furnace, converting the carbon back again to carbon dioxide.

SOLID CARBON DIOXIDE—DRY ICE

Carbon dioxide is a gas under ordinary conditions, but when cooled sufficiently, becomes solid. By compressing the gas, it is possible also to obtain liquid carbon dioxide, but the liquid cannot exist at the pressure of the atmosphere, and when allowed to escape from a tank under pressure, it evaporates so rapidly that it absorbs enough heat to freeze some of it to the solid state. This is the usual method of making solid carbon dioxide.

Since when the solid absorbs heat it passes directly into the form of gas without first melting to a liquid, it is commonly called *dry ice*. Large quantities of it are now made for use as a refrigerant in the packing of ice cream and other foods as well as for refrigerator cars. The carbon dioxide for making it is often obtained as a by-product from lime-kilns or from fermentation vats.

Its rapid evaporation keeps its temperature low, around —75°C. (—103°F.). It is thus often useful in the laboratory when temperatures lower than that of an ice and salt bath are desired.

To perform the experiments of this section you will need some dry ice. This may usually be obtained from a soda fountain or ice cream manufacturing plant, but since it evaporates rather rapidly, you had better not get it until you are ready to begin the experiments.

CAUTION: Dry ice can be handled safely with the bare hands so long as it is not squeezed. When pressed against the skin, however, heat is removed so rapidly from the area in contact with the dry ice, that a frostbite is produced which feels very much like a burn. Hence it is best to handle this material with a pair of pliers or when wearing a leather glove.

EXPERIMENT 109—Testing Dry Ice for Carbon Dioxide

Operation: Stir 2 measures of Calcium Oxide in half a glass of water for a minute or two and then let the undissolved solid settle out. Pour off the clear lime water into another glass, and drop into it a small piece of solid carbon dioxide.

Observation: At once the lime water begins to become turbid, because of the calcium carbonate formed. This, you will remember, is the usual test for carbon dioxide. If enough carbon dioxide is present, the white precipitate of calcium carbonate may later disappear, because an excess of carbon dioxide has the power to convert it into soluble calcium bicarbonate.

MEDICAL USES OF CARBON DIOXIDE

Ordinarily we think of carbon dioxide as being associated with fire extinguishers, and for this reason it is sometimes hard for most of us to appreciate that carbon dioxide can play an important part as a lifesaver and alleviator of pain. We now know that carbon dioxide stimulates respiration and serves as a means of increasing the rate of breathing. Advantage is taken of this effect in certain cases where breathing is suspended, such as during asphyxiation, physical shock, and partial drowning. Machines for administering mixtures of oxygen and carbon dioxide have proved more capable of giving relief in these types of cases than those which use oxygen alone. The breath stimulating effect of carbon dioxide, and the ventilation produced by oxygen has saved many lives. Carbon dioxide as an aid in the removal of the anesthetic when it is no longer needed, is a general hospital practice.

EXPERIMENT 110—Nickel Carbonate

Equipment: Nickel ammonium sulphate, sodium carbonate, test tubes, measure.

Operation: Dissolve one measure of nickel ammonium sulphate in a test tube half full of water. Dissolve one measure of sodium carbonate in another test tube ¼ full of water.

Result: Now add the sodium carbonate solution a little at a time to the solution of nickel ammonium sulphate, and a thick light-green precipitate of nickel carbonate will be formed.

CARBON MONOXIDE

While carbon dioxide is a harmless gas, this member of the carbon family is a violent poison. It is formed when carbon is burned with a diminished supply of oxygen. The gas burns with a blue flame, being converted into carbon dioxide.

Several thousand people are killed each year by carbon monoxide gas, which constitutes a small proportion of the gases expelled through the exhaust pipe of an automobile.

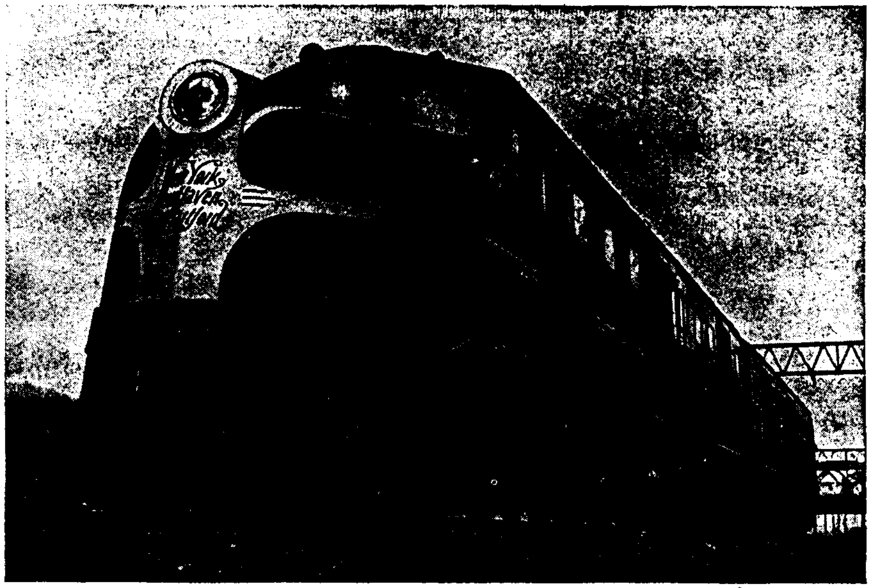
ETHYL GAS

Every automobile driver is familiar today with the trade term—"Ethyl Gas." This is an organic compound containing lead which bestows on gasoline the favorable properties characteristic of this reagent (tetraethyl lead). It is a practical anti-knock substance. When ethyl gasoline is used, it tones down the rate of combustion in the cylinders of gas engines, and pushes the piston more gently than in the case when ordinary gasoline is used. Tetraethyl lead is a dangerous substance, and warnings, therefore, accompany its use. At every filling station there are warnings posted to the effect that ethyl gas should not be spilled upon the hands or used for cleaning purposes. If ordinary precautions are followed, tetraethyl gasoline presents little danger during use.

POWER

Petroleum is today one of our chief sources of energy, and it is the chemist, who has contributed to its utility as a power agent, by the application of exhaustive research studies.

Petroleum products constituted one-half of all the supplies shipped to the American forces overseas from this country during the World War II. The consumption of petroleum-gasoline for power is staggering. For example:—Every one hundred miles traveled by an armored tank division required twenty-five thousand gallons of gasoline. And to illustrate further the magnitude of the power question the following information is very impressive: It is said that one thousand four-motored army bombers on a single six-hour flight, used about one million eight hundred and fifty thousand (1,850,000) gallons of the petroleum fuel—gasoline.



This 4,000 horsepower Diesel-Electric locomotive operates on low grade fuel oil to generate electric operating power. Unique feature of this modern locomotive is that it can pull mile long freight trains and fast passenger trains equally as well. It will pull 14 passenger coachs at speeds better than 80 miles an hour.

OIL BURNING ENGINES OF THE DIESEL TYPE

Diesel engines are now used for high-speed streamlined trains and for many passenger ships, and smaller models are being developed for automobiles and airplanes. The big advantage of Diesel engines is that instead of gasoline, they use heavy, comparatively non-volatile oil. They are therefore much cheaper to operate than ordinary gasoline engines.

Since their fuel cannot be vaporized in a carburetor and sucked into the cylinder, it is sprayed in under pressure. The hot, compressed gases in the cylinder cause it to ignite, so that no spark plug is needed. The Diesel engine has many advantages over the gas engine and great improvements have been made during the war period in its operation.

FLASH POINT

One of the things which oil chemists and engineers want to know about oils is the temperature at which they will take fire. If this temperature, called the *flash point*, is too low, there is danger of fire or explosion when the oil is used for lubricating machinery. The flash point of kerosene is important when it is used in a lamp or a stove.

To determine the flash point of an oil, it is heated slowly in a metal cup, and at intervals a small flame is passed across it at a certain height above its surface. Its temperature is read on a thermometer when it first takes fire with a flash as the flame is passed over it.

EXPERIMENT 111—Flash Point of an Oil

Equipment: Lubricating oil, spoon, alcohol lamp or candle.

Operation: Pour a few drops of lubricating oil into your spoon and warm it slowly over a flame. Soon fumes will begin to be evolved. Test them from time to time by bringing the edge of the spoon into the flame.

Result: When the flash point of the oil is reached, the fumes will take fire. Try this experiment with automobile crankcase oil, with light machine oil, and with a few drops of kerosene, note the difference between them.

EXPERIMENT 112—Preparation of Carbon Dioxide—(Effervescence)

Equipment: Vinegar, sodium bicarbonate, test tube.

Operation: Obtain some very strong cider vinegar, the stronger the better, and place about 3cc. in a test tube. Then drop into the vinegar a small measure of baking soda.

Result: Notice the violent effervescence due to the chemical action and liberation of carbon dioxide gas. You will notice that it has no odor. This is the same gas that you see bubbling out of soda-water.

EXPERIMENT 113—Vinegar and Oyster Shells

Operation: Repeat the above experiment using a piece of oyster shell. Pulverize the shell in a mortar before adding it to the vinegar. Warm the solution and test the gas given off.

Result: Oyster shells are a stable form of calcium carbonate and are decomposed by action of acids.

EXPERIMENT 114—Vinegar and Painter's Whiting

Repeat the above experiment using some painter's whiting. What is the composition of ordinary whiting?

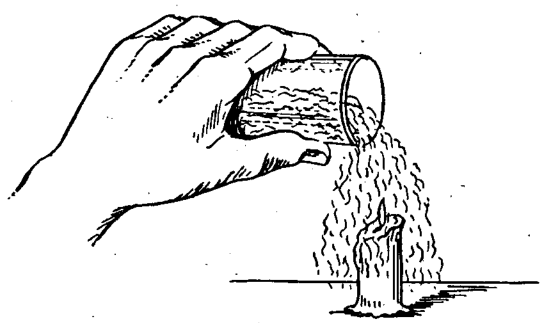
EXPERIMENT 115—Vinegar and Chalk

Repeat the above using some powdered chalk or crayon from your school room blackboard.

EXPERIMENT 116—Carbon Dioxide is Heavier than Air, and Will not Burn

Equipment: Candle, sodium bicarbonate, tartaric acid, glass.

Operation: Light a candle and set it firmly on a board by sticking it to a little melted wax from the flame of the candle. Put one half teaspoonful of sodium bicarbonate (common baking soda) in a glass and add some vinegar or a solution of tartaric acid to the glass. A violent reaction takes place with evolution of carbon



dioxide gas. Now pour the gas in the glass on the flame just as though you were pouring water out of the glass, taking care not to spill any of the acid out of the glass.

Result: Notice that the flame goes out, proving that carbon dioxide is heavier than air and will settle to the earth and also that it will not burn.

EXPERIMENT 117—Carbon Dioxide from a Burning Candle

Equipment: Calcium oxide, candle, test tubes, measure, wide mouthed bottle.

Operation: Make some lime water by putting two measures of calcium oxide in a test tube half full of water and shaking well for three or four minutes. Allow this solution to stand until clear, then pour the clear liquid into another test tube. You now have a clear solution of lime water or calcium hydroxide. Now hold a wide mouthed bottle or fruit jar over a candle flame as shown at right so that the burning gases from the flame may enter the mouth of the bottle. After allowing the gases to enter the bottle for about a minute, close the mouth of the bottle with the palm of the hand, and, inverting the bottle pour the lime water into it. Again put the palm over the top of bottle and shake for a moment.

Result: Notice that the lime water becomes turbid or milky. This turbidity is due to a white precipitate of calcium carbonate formed by the action of carbon dioxide on calcium hydroxide.

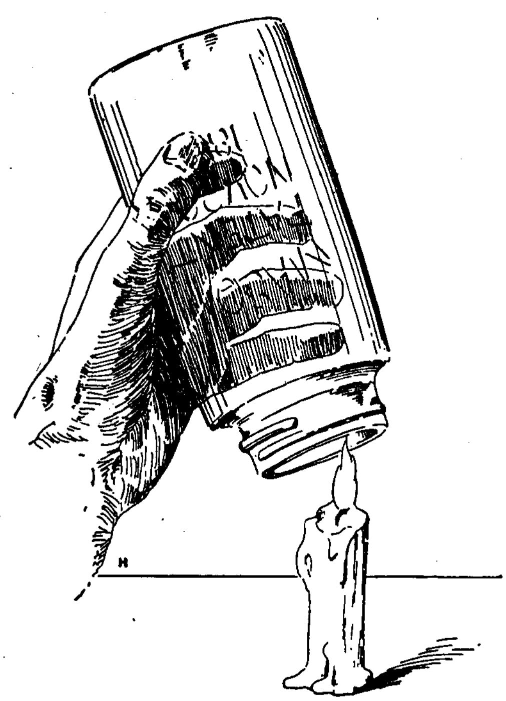
EXPERIMENT 118—The Structure of Flame—Manufacturing Lampblack

Explanation: The second or bright yellow zone of the flame contains particles of carbon heated to a white heat so that they glow brightly. The carbon is formed by the action of the heat on the gas of the inner cone.

Operation: The presence of this carbon can be shown by holding a cold spoon or piece of glass tubing in the flame for about a minute. You will notice that when you take it out it is covered with a black deposit of lampblack or soot which is one form of carbon. The cold spoon chills the flame and prevents the carbon from being completely burned.

Lampblack is made on a large scale in just this way except that natural gas is burned instead of candles and the cooling is done by means of iron pipes with water circulating through them.

Explanation: The third or outside cone of the flame consists of the gases formed by the complete burning of the carbon particles. If you will hold the cold spoon in the outer cone you will find that it gets very hot but that no soot or only a very small amount will be deposited. This cone is above the luminous one.



**EXPERIMENT 119—The Structure of a Flame—
A Gas Factory**

Explanation: If you will look closely at a candle flame you will see that it consists of three parts.

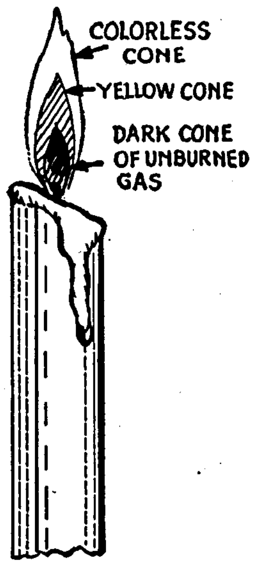
First, A dark zone just around the wick.

Second, A bright yellow zone which gives the light.

Third, A transparent zone of heat around the outside.

The first or inner zone consists of unburned gas given off from the wick of the candle. The melted grease is drawn up by a capillary action into the wick and it is there converted into gas by the heat of the flame. With care a portion of this gas can be drawn off through a tube.

Operation: Hold one end of the glass tube in the flame and directly over the wick. Hold the tube slanting upwards so that the other end is out at the side and a little above the flame. If held correctly smoke will come from the end of the tube and can be lighted with a match.



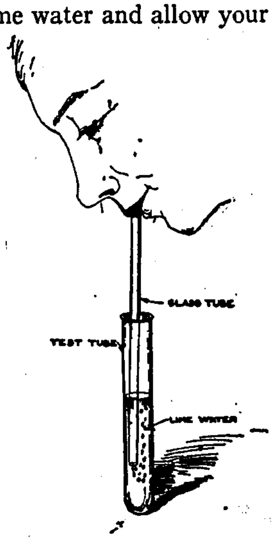
That it is relatively cool inside of the flame can be shown by thrusting a match stick into this zone for a few seconds. The portion of the stick which was held in the dark zone will not be burned as soon as that portion passing through the sides of the flame.

EXPERIMENT 120—Carbon Dioxide in the Breath

Equipment: Lime water, test tube, glass tube.

Operation: Make up a solution of lime water as previously directed. Now take a glass tube, put one end into the test tube containing the lime water and allow your breath to bubble through the lime water. Notice that very soon the water becomes turbid and after a short while a white precipitate is formed.

Result: The precipitate is calcium carbonate formed by the action of carbon dioxide in the breath upon lime water or calcium hydroxide.



PAINTS, LACQUERS, AND WATER COLORS

In all of these products you will find a solid substance furnishing body and color together with a liquid called the vehicle. It is the nature of the *vehicle* which makes the chief difference between paints, lacquers, and water colors.

In paints the *vehicle* is an oil such as linseed oil which is capable of combining with oxygen of the air to become hardened to a tough film. When such an oil has been mixed with a solid such as lead carbonate or zinc oxide it produces a good weather-resisting paint. In this case the paint will be white, but suitable coloring matters called lakes and pigments may be added to give any desired color.

If these solid colors are ground up with water containing a little glue or gum to make them stick, water colors are produced. White wash and calsomine are common examples. These are not usually weatherproof enough to be used outdoors, but are popular for inexpensive painting of inside walls of buildings.

Lacquers may be considered a kind of varnish, differing from the older types of varnishes and the paints in that the *vehicle* does not consist of an oil which dries by oxidation, but is some form of gum or resin dissolved in a mixture of solvents which evaporate to deposit the gum, together with any pigments and coloring matter, as a tough film. The lacquers may be made very quick-drying by using solvents which evaporate quickly, but if they dry too quickly the film is brittle and does not bind well to any surface it is to cover. To remedy this, a small proportion of a substance called a plasticizer is added. This is a comparatively non-volatile substance like castor oil which keeps the lacquer film tough and pliable.

Large numbers of new solvents and resins have been synthesized in the last few years, making possible lacquers of almost endless variations in properties.

**EXPERIMENT 121—Making a Water Color
Medium or Vehicle**

Equipment: Gum arabic, (purchase separately) test tubes, beaker.

Operation: Cover two measures of gum arabic in a beaker or cup with about five test tubes of water and let it stand about a day. In stirring with a stirring rod you will find that you have a thin gummy liquid or mucilage.

Application: This material when mixed with a colored powder makes a very good water color paint and after being painted on a surface, adheres when dry to the surface, and has a glossy finish. If it is to be kept long, a drop or two of a preservative such as carbolic acid must be added.

EXPERIMENT 122—A Black Charcoal Paint

Mix together three or four measures of powdered charcoal with two or three drops of gum arabic solution. This gives a black charcoal paint.

Grind small pieces of charcoal to powder on flat hard wood or metal surface, using hammer head as a pestle.

EXPERIMENT 123—A White Wash

Operation: Mix four or five measures of calcium oxide with five or six drops of water until a smooth paste is obtained. (To be sure that the lime is completely slaked with formation of calcium hydroxide.) This may be used in this form as a cheap white wash, but it is improved by addition of a binder such as gum arabic or glue solution.

EXPERIMENT 124—Making a Lacquer Vehicle

Equipment: Acetone (purchased at drug or paint store), scrap celluloid or photo film, 2 small bottles.

Operation: Obtain some scraps of clear celluloid or old photo film. If film is used, the gelatin coating must first be removed by placing the film in very hot water containing sodium carbonate. When clean and dry, cut the celluloid up in narrow strips, fill a small bottle about ¼ full of these strips, then fill the bottle nearly full of acetone. Put a stopper in and let stand for a while, shake occasionally until celluloid is dissolved.

Result: When the celluloid has gone into solution with the acetone pour the clear liquid into another bottle leaving behind any sediment. If too thick, this solution may be thinned, with acetone at any time. Add as a plasticizer, a few drops of castor oil.

You now have a very useful clear lacquer, which you may use to protect bottle labels in your laboratory, to coat polished metals to keep them bright, etc., etc. It may also be used, mixed with colored powders and lakes to make beautiful and useful colored lacquers.

EXPERIMENT 125—An Aluminum Lacquer

Operations: Moisten a few measures of fine aluminum powder with the lacquer vehicle made from celluloid. You will obtain a water-proof aluminum lacquer.

EXPERIMENT 126—A Black Lacquer

Operation: Pulverize a few measures of charcoal to a very fine powder, or obtain some in the form of lampblack. Before trying to mix it with the clear lacquer, it will help if you stir in a drop or two of acetone—just enough to barely moisten it. Now stir it up with enough clear lacquer from celluloid to make it thin enough to spread with a brush.

EXPERIMENT 127—Testing Water Color Paints for Carbonates

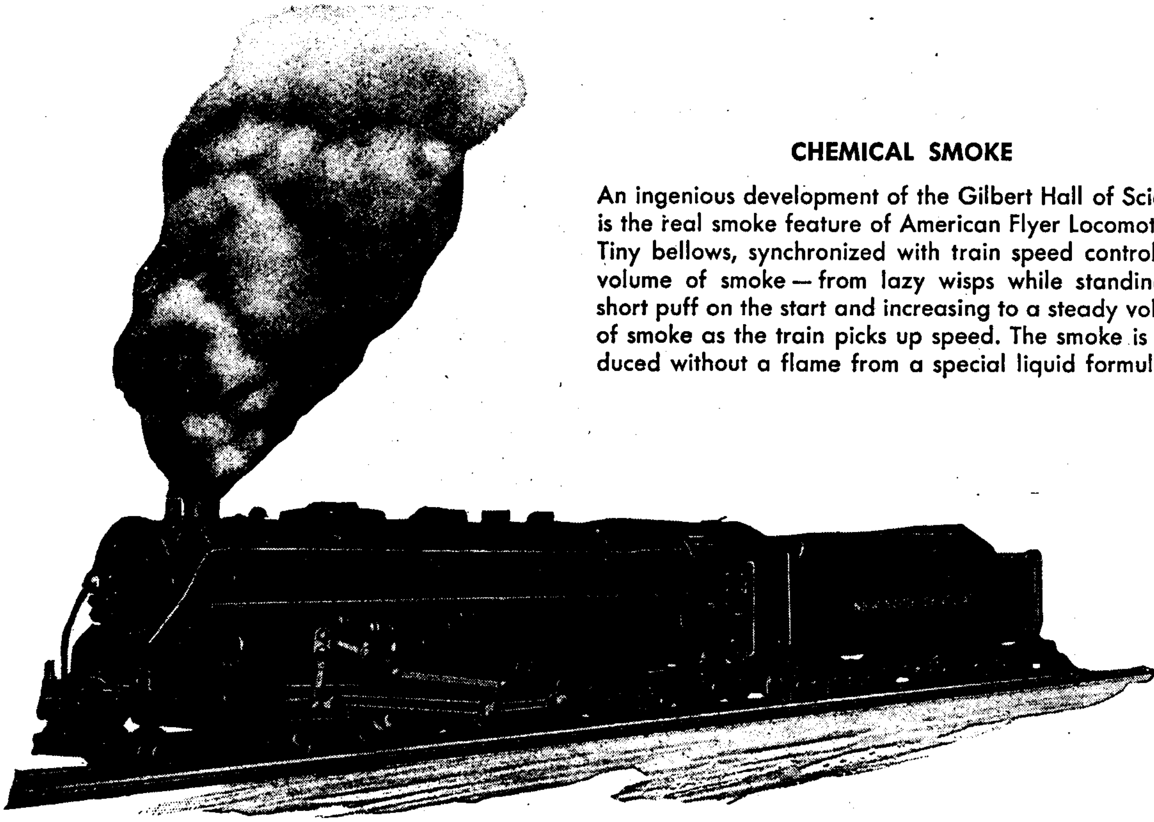
Equipment: Tartaric acid, alcohol lamp or candle, spoon.

Operation: Evaporate some of the paint in a spoon, heating until the residue is white. Allow the residue to cool, then add a few drops of tartaric acid solution or vinegar and notice the residue gives off bubbles of gas.

Result: The paint contains a carbonate, a gas (CO_2) is generated.

CHEMICAL SMOKE

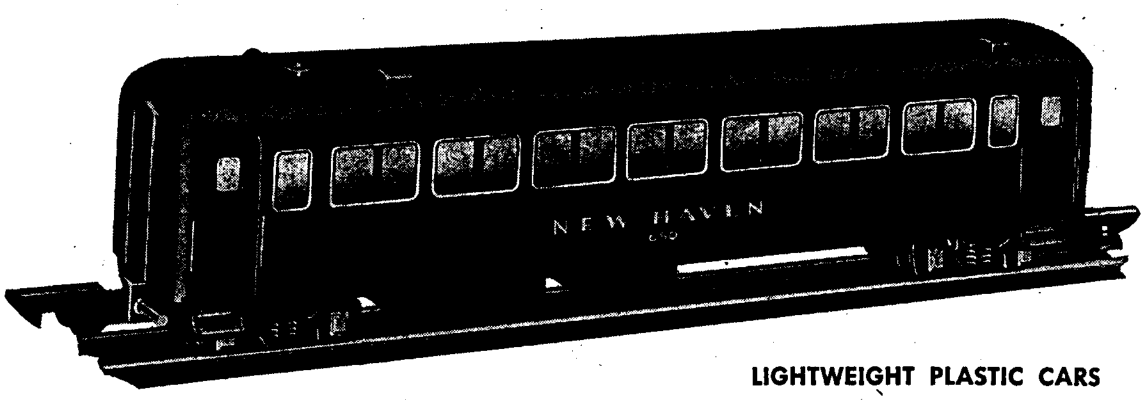
An ingenious development of the Gilbert Hall of Science is the real smoke feature of American Flyer Locomotives. Tiny bellows, synchronized with train speed control the volume of smoke — from lazy wisps while standing to short puff on the start and increasing to a steady volume of smoke as the train picks up speed. The smoke is produced without a flame from a special liquid formula.



DIECAST LOCOMOTIVE BODIES

Cast in one piece from zinc alloy — adds strength, simplicity of construction, reproduces every tiny detail and allows ample weight for good traction. Within this finely engineered body is the famous Gilbert developed electric motor that operated control valves on Grumman Hellcat planes during the war.

CHEMISTRY AIDS DEVELOPMENT OF MODEL TRAINS



LIGHTWEIGHT PLASTIC CARS

Exact 3/16" scale model cars are now made of a specially developed plastic which adds strength and yet weighs only 1/3 as much as metal cars. This enables faster pick-up speed and greater train length. Plastic never-fade brilliant colors are free from chipping and rusting — always look new. Plastic wheels give smooth operation with less noise adding a final touch of realism. These truly fine cars are a direct result of Gilbert Pioneering in plastic research at the Hall of Science.

PART XI

CHEMISTRY IN THE HOME

HOUSEHOLD CHEMISTRY

How many boys and girls have ever been instructed regarding the application of chemistry in the home? In fact, how few of us there are who really realize what an important part chemistry plays in many of the natural processes taking place in the home. Most of us are inclined to work mechanically, and do this or that according to specific directions, and how seldom do we stop to consider what is taking place chemically in the environment of home surroundings. A household kitchen is a good example of a practical chemical laboratory. It is very important today that we all realize the relationship of this household unit to our everyday life, especially simple applications of chemistry when dealing with the necessities of life, such as pure food and water, proper sanitation and clothing.

Kitchen Chemistry includes a formula, ingredients, measures, mixing bowls, thermometer, agitation, heat and refrigeration.

THE TRADE NAMES AND CHEMICAL NAMES OF 76 CHEMICALS ABOUT THE HOUSE

Trade Name	Chemical Name
Alum	potassium aluminum sulphate
Aspirin	acetyl-salicylic-acid
Bakelite	formaldehyde-phenol (resin)
Baking Soda	sodium bicarbonate
Beet Sugar	sucrose
Benzine	gasoline or petrol
Benzol	benzene
Bleaching powder	Calcium chlorohypo-chlorite
Blue vitriol	copper sulphate
Bone black	animal charcoal
Boracic acid	boric acid
Borax	sodium tetraborate
Brimstone	sulphur
Calomel	mercurous chloride
Camphor	pinene hydrochloride
Carbolic acid	phenol
Carborundum	silicon carbide
Chalk	calcium carbonate
China Clay	aluminum silicate
Common salt	sodium chloride
Corn sugar	glucose
Corundum	aluminum oxide
Cream Tartar	potassium hydrogen tartrate
Dextrose	glucose (corn sugar)
Epsom salts	magnesium sulphate
Formalin	40% solution formaldehyde
Fruit sugar	fructose
Fusel oil	mixed amyl alcohols
Gasoline	benzine or petrol
Galena	natural lead sulphide
Glauber salts	sodium sulphate
Glycerine	glycerol
Grain alcohol	ethyl alcohol
Grape sugar	glucose
Gypsum	calcium sulphate
Hypo-	sodium thiosulphate
Lamp-black	impure carbon
Lanolin	cholesterol
Lime	calcium oxide
Litharge	lead monoxide
Lunar caustic	silver nitrate
Lysol	cresol soap solution
Magnesia	magnesium oxide
Marble	calicum carbonate
Marsh gas	methane
Milk sugar	lactose
Muriatic acid	hydrochloric acid
Niter	potassium nitrate
Vitriol oil	conc. sulphuric acid
Paris green	copper aceto-arsenite
Plaster Paris	calcium sulphate
Precipitated chalk	calcium carbonate
Quick-lime	calcium oxide
Quick-silver	mercury
Red lead	lead tetraoxide
Rochelle salt	pot. sodium tartrate
Rock salt	sodium chloride
Saccharin	benzoic sulphimide
Sal ammoniac	ammonium chloride
Saltpeter	potassium nitrate
Silica	silicon dioxide
Slaked lime	calcium hydroxide
Soda (washing)	sodium carbonate
Soft soap	potash soap
Soluble glass	sodium silicate
Spirits of Hartshorn	ammonia solution
Sugar of lead	lead acetate
Sugar of milk	lactose
Table salt	sodium chloride
Talc	hy-magnesium latrate
Turnbull's blue	ferrous ferriccyanide
Venetian red	ferric oxide
Water glass	sodium silicate
White lead	basic lead carbonate
Whiting	calcium carbonate
Wood alcohol	methyl alcohol

CHEMISTRY AND PUBLIC HEALTH

Chemistry has done much to help the family doctor in his work. Today he is dependent on drugs manufactured and synthesized by chemists for the successful practice of his profession. This relationship of the doctor and the chemist is illustrated in the following table

RELATION OF CHEMISTRY TO DISEASE

Natural Drugs
The chemist has improved on the methods of isolating drugs from plants, thereby providing drugs for the physician to use.

Synthetic Drugs
The chemist has duplicated and in many cases has improved upon drugs found in nature.

Drug Adulteration
The chemist has devised methods for detecting drug adulteration, thereby protecting the public from fraud.

Bacteria and Microbes
The chemist has added to our knowledge of the life processes of these lower organisms. The physician has been taught how to prevent the destructive diseases caused by these lower organisms.

Germicides and Antiseptics
The chemist has prepared valuable antiseptics for the use of surgeons which prevent danger from infection after serious operations.

Anesthetics and Saporifics
The chemist has enabled the surgeon and dentist to perform painless operations by providing many new organic combinations possessing anesthetic properties.

Light Radiation
The chemist has cooperated in making effective the application of artificial radiation. As a result the physician is enabled to control the ravages of specific diseases and regulate proper nutrition for growing children.

Radium Extraction
The chemist has worked out practical methods for extraction of radium from ores, thereby providing an effective method of treating cancer.

Metallurgy of Iron and Steel
The chemist has provided better instruments to be used by the surgeon by making new alloys of iron and developing improved methods for tempering steel.

X-Ray and Fluoroscope
The chemist has synthesized several new chemicals for X-ray plates and fluoroscope screens. These enable physicians to make more accurate diagnosis.

FIRST AID DON'TS

Don't touch a wound with your fingers or any instrument.

Don't put an unclean dressing or cloth over a wound.

Don't allow bleeding to go unchecked.

Don't move a patient unnecessarily.

Don't allow a patient with a fracture or suspected fracture to be moved until splints have been applied.

Don't neglect shock.

Don't burn a patient with an unwrapped hot-water bottle or other heated object.

Don't fail to give artificial respiration when needed.

Don't fail to remove false teeth, tobacco, and chewing gum from the mouth of an unconscious person.

Don't permit air to reach a burned surface.

Don't wash wounds.

Don't reduce dislocations, except of the finger and lower jaw.

Don't put a quid of tobacco on a wound.

Don't leave a tourniquet on over 20 minutes without loosening.

Don't forget to send immediately for a physician.

Don't forget to determine immediately the name of the wounded.

FIRST AID CHEMICALS

Some common drugs and chemicals which are safe to apply externally for local treatment to the skin and wounds in case of minor injuries:

CAUSTICS

Silver nitrate solution	Sodium carbonate solution
Tincture of iodine	Zinc chloride solution
Carbolic acid solution (Phenol)	Sodium hydroxide solution
	Potassium carbonate solution

ANTISEPTICS AND DISINFECTANTS

Boric acid	Hydrogen peroxide
Salicylic acid	Potassium permanganate
Creosol solution	Thymol
Phenol or carbolic acid solution	Chlorinated lime
Formaldehyde	Iodoform
Hexylresorcinol in form of St. 37	Sulphur

STYPTICS

Ferric chloride

Alum (ammonium or potassium aluminum sulphate)

ASTRINGENTS

Tannic acid	Silver nitrate
Alcohol	Bismuth subgallate
Alum	Ferric sulphate

PROTECTIVES (SURFACE TREATMENT)

Zinc oxide	Magnesium carbonate
Talcum powder	Bismuth subnitrate
Collodion	

EMOLLIENTS

Paraffin	Fixed oils and fats
Petrolatum	Lanolin

LOCAL ANALGESICS FOR PAIN AND ITCHING

Aqueous ammonia	Sodium bicarbonate
Carbolic acid solution	

EMERGENCY MEDICINES

Ammonia vapor: Inhaled through the nostrils is a stimulant and delays fainting. Do not breathe strong ammonia vapors.

Tincture of arnica: Useful as a liniment and valuable as a liniment for treatment of sprains and bruises. Almost always found in mother's medicine cabinet.

Bicarbonate of soda: Known as baking soda. Use in treatment of burns.

Camphor: Comes in gum form and also as spirits of camphor. Camphor should not be applied directly to open wounds.

Ginger: Tincture of ginger is useful for bowel complaints. Trust to your mother to properly apply this remedy.

Glycerine: Recommended for burns. Mixed with rose water it makes a practical skin lotion.

Peppermint: Tincture of peppermint is a well-known remedy often times used by your mother.

Witch hazel: Extract of the plant, witch hazel. A remedy for sprains, contusions, wounds and swellings. Also good for chapped hands, burns, scalds and abrasions.

Vaseline: Recommended for burns and scalds. Should be in every medicine cabinet.

Carron oil: For burns and scalds. *Formula*: Mix equal parts lime water and raw linseed oil and shake thoroughly. This forms a creamy emulsion and can be used freely without harm. Olive oil or cotton seed oil may be substituted for linseed oil.

Limewater: Always available at a drug store for both internal and external use.

HELP IN CASE OF ACCIDENTS

Burns and scalds: Cover the surface of the wound with bicarbonate of soda and lay over it a wet cloth. Olive oil and linseed oil alone or mixed with limewater are useful.

Snake bite: Suck the blood from the wound and cauterize with a caustic as tincture of iodine, silver nitrate solution or sodium hydroxide.

Dog bite: Treat the same as in the case of snake bites.

Bee stings: Apply diluted ammonia solution, salt water solution, or tincture of iodine.

Fainting: Lay the patient on the back. Furnish with plenty of fresh air and moisten face with cold water. Keep head lower than rest of body.

Foreign substance in eye: Pull the upper lid down away from the eyeball over the lower lid, then release.

PHARMACEUTICAL PREPARATIONS

Capsules: Are made of gelatin often times mixed with a little glycerine. These are used for oils and solutions of drugs in oils. Hard gelatin capsules are used for dispensing powders.

Collodions: Solutions of pyroxylin (gun cotton) in mixtures of ether and alcohol or acetone. They are used for external applications.

Decoctions: These are aqueous preparations made by boiling vegetable substances in water.

Emulsions: Are aqueous preparations in which oils or resins are suspended by means of collodial substances. Fresh milk is an emulsion.

Extracts: Are preparations made by evaporating solutions of the soluble constituents of vegetable or animal matter.

EXPERIMENT 128—Preservation of Animal Matter

Method: Place a small piece of beefsteak in a large mouth bottle and add one spoonful of formaldehyde solution, which can be obtained in any drug store. Cork quickly and seal the bottle with a coating of collodion.

EXPERIMENT 129—Putrefaction of Meat

Method: Repeat the previous experiment, using only the meat, and do not seal the bottle with collodion. Let both bottles stand around at ordinary temperature. Which sample will first show evidence of putrefaction? The putrefaction can be detected by the smell.

EXPERIMENT 130—Decoction of Bayberry Leaves

Collect some bayberry leaves, dry them, and prepare a strong decoction. Filter and allow to stand. Notice the odor of the solution.

EXPERIMENT 131—Balsam

Repeat previous experiment, using balsam twigs.

CHRISTMAS EXPERIMENTS AT CHRISTMAS TIME

"A bayberry candle burned to the socket brings health to the home, food to the larder, and gold to the pocket."

EXPERIMENT 132—Making a Bayberry Candle

Directions: Select a piece of glass tubing of proper diameter and length for moulding a candle. Insert in one end of the tube a cork through which is drawn a string to serve as a wick. Hold the wick string taut, fill the tube with melted wax, and let cool. After complete solidification remove the cork and gently warm the tube. The candle will easily slip from the tube after heating.

EXPERIMENT 133—Balsam Pillows

The balsam tree is well known for its fragrant aroma. Secure some green boughs of a balsam tree and chop the twigs into small pieces. Balsam twigs in this form are very fragrant and are an excellent material for filling pillows at Christmas time.

EXPERIMENT 134—Making a Partridge Berry Bowl

Directions: If you are a wild flower lover and are interested in making a unique home decoration at Christmas time, construct a partridge berry bowl. The leaf of the partridge berry is evergreen and the berries a brilliant red. Select uniform sprays of the plant and arrange in the bowl so that both leaves and berries are visible. If properly packed the sprays will continue to grow. The partridge berry bowl requires no attention except a few drops of water once a month. Keep tightly covered to conserve water.

EXPERIMENT 135—Evergreen Trees

Every boy and girl should be familiar with our native evergreen trees. Collect sprays of the following trees: Arbor vitae, American yew, balsam fir, hemlock, juniper, red cedar, white spruce, American arbor vitae, white pine, red spruce, and Norway spruce. Compare the foliage of these species and note any characteristic differences. Which species are commonly used for Christmas trees?

AROUND THE FIREPLACE AT CHRISTMAS

EXPERIMENT 136—The fireplace is the center of family life at Christmas time, as well as in the long winter evenings. What can you suggest in the way of a Christmas decoration to make it more attractive? Make the proper selection of evergreen, etc., for roping, sprays and provide balsam needles and selected cones for burning in the fireplace.

COLORED FLAMES FOR FIREPLACE AND OPEN FIRES

Material for preparing interesting Christmas gifts for your friends can easily be made by chemically treating wood and other combustible material which when thrown on burning fires will produce colored flames. Small pieces of well dried wood, pine cones, corn cobs, charcoal, knots and rolls of old newspapers are all suitable materials for use in making colored fires.

GENERAL DIRECTIONS FOR CHEMICAL TREATMENT

To apply the chemicals, take a small wooden pail, a tub or an old earthen crock. Do not use a metal container because the chemicals will attack and destroy the metal. Dissolve the chemical to be used in the proportion of one pound to a gallon of water. In some cases it will not be necessary to use so strong a solution. After the chemical solution is ready, then take your material to be soaked (pieces of wood or cones) and suspend them in a porous bag, a wire basket or some similar container, and dip them into the chemical solution. Let them soak for two or three hours so that the solution will completely penetrate the combustible material. Then remove from the solution and allow to drain, finally spreading on paper for complete drying.

Caution: When preparing the chemical solutions and also when soaking wood and cones do not put your hands in the solutions as they are corrosive and will injure the skin.

EXPERIMENT 137—A Blue Fireplace Flame

Directions: Use blue vitriol or copper sulphate as your chemical and dissolve as directed in water. This chemical is not expensive and can be purchased at a wholesale drug house. Wood shavings, cones, coarse charcoal, and dried branches of evergreen trees make suitable material for soaking with this chemical.

EXPERIMENT 138—A Purple Fireplace Flame

Potassium permanganate is a practical reagent for cone or wood soaking. A beautiful purple color characteristic of potassium will be produced.

EXPERIMENT 139—A Green Fireplace Flame

For producing this color use a solution of boric acid.

EXPERIMENT 140—A Red Fireplace Flame

Two chemicals can be utilized for producing a red flame; lithium chloride or strontium nitrate. It will not be necessary to use more than one-quarter to one-half of the quantities of chemicals recommended in the general directions. The red colors are more intense and therefore less chemicals are required to produce the desired results.

EXPERIMENT 141—An Orange Fireplace Flame

For producing an orange fire, calcium chloride is recommended. This is a very cheap salt.

EXPERIMENT 142—Christmas Souvenirs

Directions: Prepare small bags of netting and then place some dried wood shavings or small cones soaked with different salts in each of them and tie with bright red ribbons or cord. These bags make excellent gifts at Christmas time and give much pleasure when burned in fireplaces or open hearth fires, giving colored flames.

EXPERIMENT 143—A Merry Christmas Greeting

Operation: Make a strong water solution of potassium nitrate. Using this solution as an invisible ink, write on a sheet of white, unglazed paper the following: "Merry Xmas." Make the lines heavy and writing continuous without any breaks. Allow the paper to dry thoroughly.

Result: Touch the beginning of the writing at the letter "M" with a red hot wire. A spark will develop and run over the paper tracing out the words "Merry Xmas." This is an example of fire writing.

EXPERIMENT 144—Happy New Year

Repeat the above experiment using the greeting "Happy New Year." Make writing continuous.

EXPERIMENT 145—April Fool

Repeat the above experiment, using the words "April Fool."

EXPERIMENT 146—Believe It or Not

Repeat the above experiment using the well-known Ripley phrase "Believe it or Not."

THE CHEMISTRY OF FOODS

While there is quite a variety of foods and the compounds present in foods are very numerous and often very complex, yet they may all be included in a few general classes.

The edible portion of our foods consists essentially of proteins, fats, carbohydrates, mineral matter and water.

Proteins, which occur to a larger extent in animal foods serve to replace the worn-out tissues of our bodies and to supply material for growth. The carbohydrates and fats are both oxidized in the body to carbon dioxide and water and consequently serve as a source of heat and muscular energy. If carbohydrates or fats are lacking in the foods we eat, the protein material in them furnishes the necessary heat and energy. The mineral matter supplies the necessary material for building up the solid tissues of the body besides taking other complex parts. Proteins occur in both animal and vegetable foods and are composed of the elements carbon, hydrogen, oxygen, nitrogen and small amounts of sulphur and phosphorous. Carbohydrates are chiefly found in vegetable foods in the form of starch and sugars and are composed of the elements, carbon, hydrogen and oxygen. Fats occur in both classes of foods, but occur in much larger quantities in animal foods. They are compounds of glycerine with organic acids and are composed of the elements, carbon, hydrogen and oxygen.

A Balanced Diet for Health

It is important, since the various constituents of our foods serve different purposes, that we use the proper proportions of these materials in order to keep up a healthy body. For example, during the winter months we should eat a larger amount of animal foods since they contain a larger proportion of fats necessary to keep the body warm. During the warm weather we should eat less meats or animal foods and more vegetables or fruits. They contain carbohydrates which serve to keep up the necessary body heat and energy. How often people are made very uncomfortable and sometimes sick by their ignorance of the amounts and kinds of foods that the body requires during the different periods of the year. Did you ever realize that sometimes a heavy cold is brought about by over-eating? This happens quite often when we are not exercising the body enough; the foods are not oxidized or burned up, the blood becomes stagnant and we are sick. The body is a very intricate form of machinery, each part performing its own particular function the same as the different parts of a watch. If we abuse any one of these parts by eating too much with an insufficient amount of exercise or by not eating the right kinds of foods, something goes wrong and we are sick. Throughout the country there are chemical laboratories which are conducted by the Government and local public health boards for the purpose of analyzing and testing foods as to their purity. A few years ago the Government enacted the Pure Food Law, which requires that foods which are bought by the public should be free of impurities, of adulterants as they are called. Adulterants are inferior materials sometimes mixed with foods to make them cheap. For example, oleomargarine was quite often used to take the place of good butter, whereas, it was only a substitute. In other foods harmful chemicals were often used to preserve them. Quite often foods were adulterated to give them weight, bulk or proper color. For example, coffee was often weighed with chicory. Today things are quite different; most of the foods on the market are pure or nearly so.

Proteins and Minerals

The square meal today has come to have six sides, all of which are necessary for human beings, if they wish to enjoy good health. First, water is the most important of these sides. Proteins; bones require proteins and minerals. If human beings do not get a proper amount of proteins and mineral substances, for bone growth, they suffer from rickets and other bone disorders. Calcium and phosphorus are the two chief mineral elements needed for bone growth. It is very important to realize the importance of calcium as one of the essential elements in the metabolism of human beings. In the growth and repair of the body frame-work and teeth of the body it is very necessary. Calcium plays a very distinct role in the tuberculosis patients' bid for the return of good health. A careful study of all available data indicated that about 95% of school children suffer from tooth decay. The prevention of tooth decay would be an untold boon to the public. A lack of vitamin C in the diet leads to an increase in tooth decay. The lack of this vitamin produces a degeneration and retards the nutrition of the teeth. Vitamin D is a potent substance and also a regulator of tooth decay. Blood serum high in phosphorus is essential to tooth immunity.

Carbohydrates, Body Fuel

The chief fuels used by the body are known as carbohydrates. Two forms of carbohydrates which man uses as food are starches and sugars. Foods rich in these substances are important energy-giving foods. Fats; These are important constituents of such foods as butter, cream, bacon. Fats serve two purposes in the body; they serve to keep the body warm by supplying a padding around the muscles; they also serve as a reserve fuel supply, and are drawn upon by the body as a source of energy. Materials which regulate body processes; common salt, iron, mineral substances, and the naturally occurring principals known as *vitamins.*

VITAMIN AND HORMONE SUBSTANCES

With the increased knowledge of the vitamins, hormones, and other active principals found in tissues, the organic chemist has been called upon again to determine the organic structure of these naturally occurring substances. We shall not take up space in this text to give the results of the extensive work which has been accomplished. The problems involved are highly complicated, and involve substances and techniques which cannot be presented understandingly to boys and girls for whom this book is edited. It should be made clear that the existence of these substances depends upon the development of biological tests for the entities concerned, and without such tests the organic chemist would be helpless. A fundamental distinction between vitamins and hormones is that hormones are synthesized by the body, whereas vitamins cannot be synthesized but must be furnished by the food. It should be emphasized that vitamins have no chemical characteristics in common, and neither do the hormones. Vitamins, so far as known, are relatively simple compounds, whereas, the hormones may be simple or have the complexity of proteins.

THE VITAMINS ARE CLASSIFIED AS FOLLOWS:

a. Fat soluble compounds possessing vitamin activity; The four carotenes C_{10} H_{56} found in the pigment of carrots are the source of vitamin A (caritol). B carotene is the best known of these pigments. Life and growth of higher animals are impossible without vitamin A. Sweet potatoes are a source of vitamin A.

b. *Calciferols* are closely related to the sterols and are anti-rachitic and possess vitamin D activity. Rickets is a very common disease among poorly fed children.

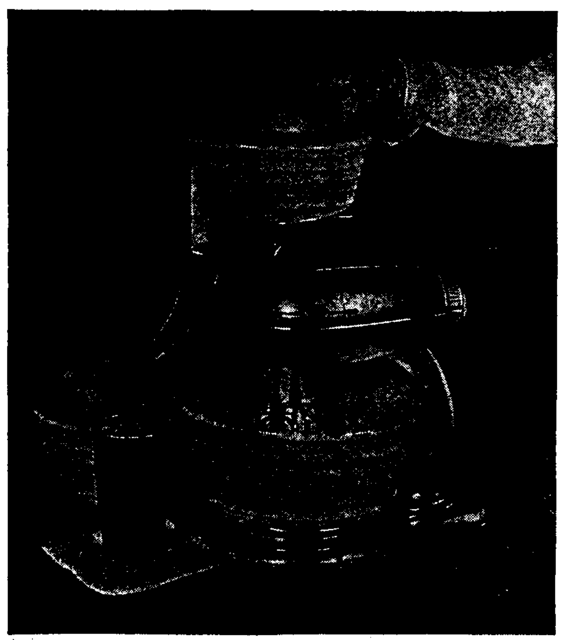
c. *Tocopherols* possesses vitamin E activity which prevents sterility.

d. *Phthiocol* and *Methylnaphthoquinone* are products possessing vitamin K activity. They have the power of preventing delayed clotting of blood in chickens.

WATER SOLUBLE COMPOUNDS POSSESSING VITAMIN ACTIVITY

1. Thiamine seems to be fundamental in all forms of life. In man its lack produces beri-beri, a polyneuritis. It is the only natural substance possessing vitamin B_1 activity.

2. Ascorbic is acid active vitamin C. Found in fruit juices, it has the power of preventing scurvy. It is the



Orange Juice, a good source of Vitamin C, easily extracted with a Gilbert Food Fixer.

most abundant of the known vitamins. Gooseberries contain as much as 2% of ascorbic acid.

3. *Riboflavin* is the water soluble, yellow pigment of milk (whey). It is an essential constituent of the diet of human beings. No other naturally occurring substance is known which can take the place of riboflavin in the diet.

4. *Nicotinamide* is a pyridine compound that has been known for years. It is a vitamin which has the power of preventing pellagra.

5. *Pyridoxin* (B) is a vitamin which will prevent "rat pellagra."

6. *Pantothenic* acid was discovered by the American chemist, R. G. Williams. It is an essential constituent of the diet of animals and fowls, and is a vitamin affecting growth and respiration of many organisms and tissues. No other naturally occurring substances can replace this compound as a food constituent.

7. *Biotin* has the physiological activity of vitamin H. Egg white contains a substance which combines with biotin and renders it available.

ORGANIC SUBSTANCES POSSESSING HORMONE ACTIVITY

a. *Carbon dioxide* possesses activity similar to that of a hormone. It is indispensable as an agent for inducing respiratory movements.

b. *Acetylcholine* is liberated by various types of nerves.

c. *Adrenaline* is found in the adrenal glands.

d. *Corticosterone* is a compound also isolated from the adrenal gland. Its physiological activity is too complex to discuss here.

SEX HORMONES

a. Various sex hormones (6) are known. Three are classed as female sex hormones, and three as male sex hormones.

b. *Thyroxine* is a derivative of the ammino acid tyrosins. It is the active principle of the thyroid gland. It is characterized by the presence of iodine.

PLANT HORMONES

a. *Ethylene* is produced by plants during the ripening of fruits.

b. *Auxin* is produced by plants and serves as a growth regulator. It also occurs in animal tissues, and urine has been found to be the best source.

c. *Traumatic acid* is a hormone produced as a result of injury to the tissues of bean pods.

HOME KITCHEN CHEMISTRY

EXPERIMENT 147—The Use and Care of Aluminum Ware

Caution: Neither strong acids nor alkalies should be used in cleaning aluminum ware. Instead lemon juice in water boiled in the vessel to be cleaned will impart a beautiful luster to the surface. This should be in the proportion of two teaspoons lemon juice to each cup of water. Boil for 10 to 12 minutes.

EXPERIMENT 148—Why Zinc Vessels Are Not Used For Food

Caution: Acids and alkalies easily dissolve zinc; therefore, foods should not be prepared in zinc vessels. Sodium chloride solution also attacks zinc readily, and as the soluble compounds of zinc are poisonous, it would be very dangerous to use zinc vessels in preparing foods.

EXPERIMENT 149—Testing Silver Plate

Equipment: Sodium bisulphate, sodium carbonate, test tube, measure, alcohol lamp or candle, gas delivery tube.

Operation: To find out whether an article is plated with nickel or silver, prepare a solution of four measures of sodium bisulphate in a test tube half full of water. Put the article which you wish to test in this solution and warm it for a few minutes.

Result: If a greenish solution is formed, the article contains nickel. Test the solution by adding sodium carbonate until it no longer effervesces and then generate hydrogen sulphide gas, and pass the gas through your delivery tube into the solution. A black precipitate indicates nickel.

EXPERIMENT 150—Cleaning Silverware

Operation: Sodium thiosulphate dissolves silver sulphide. Prepare a solution of two measures of this salt in a test tube three-quarters full of water. Moisten a soft cloth in this solution and rub a piece of silverware, which is tarnished with silver sulphide.

Result: The silver sulphide will be quickly removed.

EXPERIMENT 151—A Test for Soil in the Flower Pot

Operation: Mix a sample of the soil with water to make a thin paste. Insert a strip of blue litmus paper and allow to remain for ½ hour. Withdraw the paper and wash with water.

Result: If the paper has turned pink the soil is acid.

EXPERIMENT 152—A Fragrant Paste Wax

Operation: One part of carnauba wax, and one part of bayberry wax. Mix with sufficient turpentine to form a paste. Mix thoroughly.

Apply this mixture to a piece of furniture that your mother wants renovated. Rub briskly for a good polish.

EXPERIMENT 153—Colored Soap

Operation: Put a spoonful of shaved soap in a cup and add 2-3 drops of cochineal solution. Grind thoroughly and mix until of uniform color.

Result: What is the color of the soap? By using different dyestuffs almost any color of soap can be made.

EXPERIMENT 154—Perfumed Soap

Operation: Put one heaping spoonful of soap shavings in a mortar, and then add 2-3 drops of strong cochineal solution. Grind thoroughly and then form into a cake by pressing into a smooth mould.

Result: Various perfumes can be used to make a soap of a desired odor.

EXPERIMENT 155—Liquid Soap

Operation: Put a spoonful of soap shavings in a mortar and add water, a few drops at a time, stirring the mixture continually with the pestle. When the proper consistency has been reached pour the liquid soap into a bottle. This is a useful toilet preparation.

EXPERIMENT 156—How Soap Cleanses

Have you ever thought why it is that soap will immediately remove dirt from your hands; while if you try to wash them without soap it takes very hard rubbing?

Explanation: Soap cleanses by emulsifying or holding the dirt in suspension and then by rinsing the soap away with water, the dirt goes away with it. On the other hand, if you wash with water which does not contain soap, the water will loosen the dirt, but in order to remove the dirt it must be rubbed away. Try this experiment.

Operation: Obtain a small amount of grease or oil and fill a test tube about one-quarter full with it. Now add water until the tube is half full and shake the tube for a few minutes until the oil and water are thoroughly mixed. When you stop shaking, the oil and water separate very rapidly. Now add a small amount of soap solution to the oil and water in the test tube and shake again. This time the oil and water stay mixed together for some time before they separate, and it would be an easy matter to remove the oil while it is held by the soap and water.

EXPERIMENT 157—How Biscuits are Raised

Equipment: Sodium bicarbonate, flour, alcohol lamp or candle, measure, spoon.

Operation: Put a teaspoonful of flour in a tumbler and add four measures of sodium bicarbonate. Stir this mixture with a little water until you have a dough like bread dough. Now put a piece of this dough about the size of a marble on your spoon and heat it over a flame. Notice that the dough swells and that it becomes porous or light, due to the gas that is formed in it.

Result: Ordinary baking soda is sodium bicarbonate; when heated it gives off part of the carbon dioxide gas which it contains, so when used in biscuits, cakes, etc., the gas coming off raises the dough and makes it porous. In cooking, sour milk, or other acid, is used with the baking soda. This neutralizes the soda and gives off more carbon dioxide gas.

SOME APPLICATIONS OF CHEMISTRY IN MOTHER'S KITCHEN

All the food on your dinner table must be grown, processed, stored, shipped and packaged. In all these operations inorganic and organic chemicals work for you by protecting the food against destructive organisms. These chemicals represent various forms of insecticides and fungicides which growers use in the field, garden and orchard as spraying and dusting material. Also various fumigants are used to prevent infection in milk, warehouses and railroad cars. Here again we are made to realize the value of the chemist in devising and regulating all the processes of production and distribution.

We will now examine some products that are present in mother's kitchen:

EXPERIMENT 158—Protein In The White Of An Egg

Equipment: Calcium oxide, test tube, measure, test tube holder, alcohol lamp or candle.

Operation: Place in a test tube some white of a boiled egg and mix well with one measure of pulverized calcium oxide. Then add three or four drops of water and heat the mixture over a flame and smell the gas that is generated. Conduct some of the gas evolved over a moistened red litmus paper.

Result: Proteins contain nitrogen. Egg white is a protein substance and on heating with calcium oxide is decomposed with liberation of ammonia.

EXPERIMENT 159—Decolorizing Hamburg Steak

Meat turns a deep red color on long exposure to the air. This is due to changes in the blood content.

Operation: To bleach out meat put a small quantity of chopped meat in a test tube and add some hydrogen peroxide solution. Warm the solution gently and let stand. The chopped meat will be bleached and lose all its color. Hydrogen peroxide is an excellent preservative for animal matter and will prevent such material from putrifying.

EXPERIMENT 160—Tannin in Tea

Equipment: Ferric ammonium sulphate, tea, measure.

Operation: Make a very strong water solution of hot tea. Cool and pour off the clear liquid and add about one measure of ferric ammonium sulphate. Shake well.

Result: On standing a black color will develop due to the formation of iron tannate.

EXPERIMENT 161—A Low Temperature Bath

Follow the technique of the ice cream manufacturer and mix some crushed ice with pulverized salt. Read the temperature with a thermometer.

EXPERIMENT 162—A Sugar Bath

Repeat the preceding experiment, using some household granulated sugar in place of pulverized salt. Do the sugar and salt produce the same results? Explain.

EXPERIMENT 163—Vegetable Stains on Hands

Rub with a fresh cut slice of raw potato.

EXPERIMENT 164—Testing a Raisin for Iron

Equipment: Hydrochloric acid solution, sodium thiocyanate, potassium ferrocyanide, raisin, porcelain crucible, alcohol lamp, funnel and filter paper.

Operation: Select a raisin and wash it with water. Then incinerate the raisin in a porcelain crucible by heating the crucible over your alcohol lamp until the raisin is reduced to an ash. This ash contains only inorganic matter. Digest the ash in diluted boiling hydrochloric acid solution and filter. The insoluble residue will be silica.

Result: The iron content of the ash will be dissolved, forming iron chloride. Apply the characteristic test for iron by adding some of the iron solution to the reagent sodium thiocyanate. If iron is present red iron thiocyanate will be formed. Also add some of the iron solution to a mixture of potassium ferro and potassium ferricyanide solutions. A blue color will be formed if iron is present.

STARCH AND SUGAR INDUSTRY

Before we say anything about this industry, we will discuss briefly how starch and sugar are being made every day by nature in plant life and the changes these substances undergo in animal life.

The carbon dioxide which is always found in the air is made by the heat of the sun to combine with water in the plant tissues to produce a compound formaldehyde. Formaldehyde in turn undergoes what is known as polymerization in the plant to form sugars. By polymerization we mean several molecules of the substances combining to form a more complex compound. The sugar is changed to starch by plant ferments of enzymes. In the process of converting carbon dioxide and water by the plants into sugar and starch, oxygen is liberated and given off into the air. The process may be expressed very clearly as follows:

1.	CO_2 Carbon Dioxide	$—H_2O$ Water in presence of chlorophyll.	Sunlight	CH_2O Formaldehyde	O_2 Oxygen
2.	$6CH_2O$ Formaldehyde	$C_6H_{12}O_6$ Dextrose (sugar)	Enzyme	$C_6H_{10}O_5$ Starch	

Man is dependent upon plants for starches and sugar. When these foods are taken into the body they are changed back again, with the aid of the oxygen which we breathe into the lungs from the air, into water and carbon dioxide. The carbon dioxide is constantly being passed out of the body when we breathe the air out of our lungs.

To summarize, then, we can say that carbon dioxide of the air is taken up by plants and oxygen given off. In animal life oxygen is removed from the air and carbon dioxide given off. In other words, carbon dioxide sustains plant life while oxygen sustains animal life.

Sources of Starch

Nearly all plant and vegetable matter contains starch —potatoes, corn, wheat, and rice containing a large percentage of this substance. Sugar cane and beets on the other hand are the chief sources of sugar. Starch and sugar come under the class of compounds known as carbohydrates That is, they are compounds of carbon with hydrogen and oxygen. In the United States starch is obtained chiefly from corn, while in Europe potatoes are the chief source. Starch consists of minute granules composed of a substance known as granulose, which is surrounded by a membrane composed principally of cellulose. Starch does not dissolve in cold water because the granulose which is soluble in water is protected from the action of water by insoluble cellulose membranes. When heated with water, the granules burst and the granulose dissolves, forming the well-known starch paste.

Besides its commercial use in the laundry, starch is manufactured for the purpose of making the commercial products known as glucose, grape sugar, malt glucose, dextrin, British gums and soluble starch. These compounds are all formed from starch by hydrolysis with suitable reagents. They are all important compounds and are widely used. Corn syrup is nothing more than glucose which originally came from starch through hydrolysis.

Cane sugar or ordinary household granulated sugar is obtained from the sugar cane through a process of refining. It is the sweetest of all natural sugars and is called sucrose. It has the formula $C_{12}H_{22}O_{11}$. On hydrolysis with acids this sucrose may be broken down into two other sugars, dextrose and levulose, having the same formula $C_6H_{12}O_6$, but differing in physical properties. Levulose is sweeter than dextrose and is the chief component of honey. Dextrose is found in commercial glucose or corn syrup. Maltose is the sugar that is formed by the action of enzymes, such as diastase on starch. It has the same formula as dextrose and levulose and is the sugar from which alcohol and most of the alcoholic liquors are prepared.

SUGAR

Do you know that the pure compound sugar ranks first in the production of pure organic compounds, and is prepared in the United States in a larger quantity than any other organic compound. This is a fact not commonly realized even by many chemists. Statistics show that the amount produced is about seventy-five times the amount of its nearest contender.

The problems to be undertaken by a research program are very numerous and cover a broad field. They include studies in nutrition, bio-chemistry, physiology, medicine, dentistry, microbiology and both fundamental and applied organic chemistry.

EXPERIMENT 165—How to Make Corn Starch

Equipment: Water, test tube, mortar and pestle.

Operation: Place about 15 kernels of corn in a mortar, or cup. Add a test tube full of water and allow the corn to soften somewhat in the water for several hours. Then grind the kernels up into a pulp by means of a pestle or spoon.

Add a little more water and after stirring a few times, strain the mixture through a coarse cloth into a glass. Fill the glass half full of water, stir the mixture a few times and allow the liquid to stand quietly for a few minutes.

Result: Notice the sediment in the bottom of the glass. This is starch originally contained in the corn.

Corn starch is manufactured extensively in the United States by a method similar to this.

EXPERIMENT 166—How to Make Starch From Flour

Equipment: Flour, water, cup.

Operation: Besides starch, flour contains an organic compound called gluten. In order to separate starch from gluten proceed as follows:

Put two tablespoonfuls of flour in a cup and add just enough water to make the flour into a stiff dough. This can be done by working the moist flour with the fingers until it is kneaded into a round ball the same as your mother does when she prepares bread dough for making bread.

Result: Now take this ball of dough and holding it in a glass nearly full of water, work it with the fingers so that all portions of it will come in contact with the water. Notice that the starch separates from the dough and settles in the water leaving a very sticky mass of gluten in the fingers.

EXPERIMENT 167—Manufacture of Potato Starch

Operation: Obtain quite a large potato and by scraping reduce it to pulp. Mix this pulp with an equal amount of water and then squeeze through a cotton cloth (linen is also good). Repeat the squeezing operation several times until the pulp is quite free of all water. The pulp (which we can now call cellulose) remains on the cloth and the starch and water goes through into a container.

Result: Allow the solution of water-starch to set for some time and then pour the water off. The white compound in the bottom of the container is potato starch. By drying this starch of all water a good quality dry starch will be obtained. The principle involved in the manufacture of starch from potatoes in Europe is similar to this, although complicated machinery is used in the manufacture of large quantities and in the purification and separation of starch.

BREAD MAKING

In making bread, biscuits, cake, etc., baking powder is thoroughly mixed with the dry flour. Water is then added, the flour is kneaded into a dough and put in the oven. Heat nearly always hastens a chemical reaction; that is why the dough raises quickly in the oven.

The moisture plus the tartaric acid act upon the sodium bicarbonate and start the formation of carbon dioxide Soon the dough begins to rise, due to the gas being liberated. The powder has been distributed through the mass, so that a million gas bubbles are formed which puff out the dough, making it porous. During the process of baking, the gas and the remaining constituents of the baking powder are converted into harmless substances.

EXPERIMENT 168—How to Make Baking Powder

Equipment: Sodium bicarbonate, tartaric acid, (household cream of tartar) test tube, measure.

Operation: Put two measures of sodium bicarbonate and two measures of tartaric acid into a test tube. Mix thoroughly by closing the mouth of the tube with your thumb and shaking back and forth. Notice that nothing happens. Now add a half test tube full of water and notice the sudden reaction. These are the two active constituents of baking powder. The gas liberated is carbon dioxide.

Result: Baking powder is composed of two or sometimes more chemicals, which, when dry, are not active. When water or moisture is applied, the tartaric acid and sodium bicarbonate are put into forms whereby they react readily upon each other, the sodium bicarbonate giving up its carbon dioxide. This gas is the same as that formed in the lungs or in soda water.

BAKING POWDER

Baking powders found on the market are of three kinds, named from the kind of acid salt used. For example, tartrate powders have tartaric acid or acid potassium tartrate (cream of tartar); phosphate powders have an acid phosphate such as lime or potassium; and alum powders have as an acid aluminum sulphate.

The reactions which take place in these different baking powders may be expressed as follows:

1. Acid potassium tartrate and sodium bicarbonate—carbon dioxide and sodium potassium tartrate (Rochelle Salts).

2. Acid potassium phosphate and sodium bicarbonate—carbon dioxide and sodium potassium phosphate.

3. Aluminum sulphate and sodium bicarbonate—carbon dioxide and aluminum hydroxide and sodium sulphate (Glauber's Salt).

TESTING MILK

No food product is so susceptible to contamination as ordinary milk. Consequently extreme care should be made to protect it from conditions that lead to infection and absorption of bacteria. Clean bottles must be provided for storage and these must be very carefully sealed. Do not allow the milk to be exposed to the atmosphere, and keep well cooled by storage in a good refrigerator. Milk is one of our essential foods for the young, and every precaution should be taken to avoid contamination and bacterial infection.

EXPERIMENT 169—How to Test for Coloring Matter in Milk

Equipment: Sodium bicarbonate, milk, test tube, measure, filter paper.

Operation: Put one measure of sodium bicarbonate in a test tube one-third full of milk, and, putting the thumb over the mouth of the test tube, shake the contents of the tube thoroughly. Now place a piece of filter paper in this solution and allow the paper to remain in this solution for 10 to 12 hours. At the end of this time remove the paper and examine it. If it is colored a reddish yellow, the milk contains some artificial coloring.

EXPERIMENT 170—How to Test Butter

Equipment: Butter, spoon, alcohol lamp or candle.

Operation: 1. Place a small piece of butter about the size of a pea in your spoon and heat it slowly over a flame. If the butter foams and froths on boiling the butter is fresh. On the other hand, if it sputters or pops, it is either oleomargarine or renovated butter.

2. Another test that may be applied to butter to show whether it is adulterated is as follows:

Operation: Heat a half cupful of milk until it is very hot. Then put into the milk a piece of butter about a half inch square to be tested and stir until it is all melted. Now place the cup in a pan of cold water containing a little ice and stir the milk continually until the milk becomes cold and the butter solidifies.

Explanation: If the butter is pure or renovated it will solidify into small particles throughout the milk. If it is oleomargarine it will solidify into one solid cake.

Pure fresh butter contains water and butter fat. Butter fat consists principally of the fats olein, palmitin and stearin. The flavor of the butter is due to the presence of a small amount of butyrin, which is an ester of butyric acid and glycerine.

BUTTER TEST

There are three different kinds of butter to be taken into consideration. First is unsalted butter which is butter fat removed from whole cream by the process of churning. This is known as sweet butter. Such butter very quickly becomes rancid if not kept in a cool place. The ordinary creamery type of butter is sweet butter to which common salt has been added. The salt serves as a preservative.

The third class of butter is renovated butter. This is made from rancid butter and is accomplished by aeration of spoiled butter moistened with sweet milk. The aeration seems to remove unpleasant odors. Renovated butter is used in bakeries, but has not the flavor of fresh butter.

A valuable substitute for ordinary butter is oleomargarine. Vegetable oils and animal oils are used in cocoanut oils. It cannot be condemned as a food product. It is made under strict government supervision and under sanitary conditions. In most states we have laws regulating its manufacture, and every means is taken not to defraud the purchaser by false advertising.

EXPERIMENT 171—How to Make Casein

Equipment: Sodium bisulphate, milk, test tube, measure.

Operation: Put three measures of sodium bisulphate in a test tube one-third full of water and shake until it dissolves. Now add a few drops at a time, with constant shaking, some of this solution to a test tube half full of milk. Notice that suddenly a white curdy precipitate is formed.

Result: This precipitate is casein. This is the same precipitate that is formed when milk turns sour or curdles. The souring of milk is brought about by the formation of lactic acid, due to the fermentation of the milk.

Casein, besides being used as a food in cheese, has many other important uses today. It is used in the manufacture of adhesives, paints, in dyeing, in medicines, as electrical insulators and in making plastic masses as in stoneware, toys, etc. In the manufacture of these things the casein is put through a certain chemical treatment.

EXPERIMENT 172—Milk and the Common Cold

Fresh pure milk is a perfect food and contains necessary vitamins for promoting growth. Faithful use of this food will aid in protection against the common cold. The true cause of a cold is unknown but we do know that a diet high in vitamin A shortens a cold's duration and lessens its severity. Therefore, eat foods rich in this vitamin. Some of these foods are: Milk, butter fat, cream, cheese, eggs, liver, cod-liver, cod-liver oil, fish-liver oils, red salmon, green and yellow vegetables, fruits, tomatoes, and olives.

SOAP

Most of us are familiar with soap and its cleansing properties, but few of us understand the chemistry of soap making and the part soap plays in cleansing. Soap was made in early history by treating fat with the lye obtained by extracting wood ashes with water and lime. What really happens when ashes are treated in this way is that the potassium carbonate present is dissolved by the water. This water solution of potassium carbonate then reacts with the lime or calcium oxide to form an insoluble compound of calcium carbonate and a soluble compound of caustic potash, potassium hydroxide or lye as it is commonly called. Fats or grease, which are compounds of glycerine with fatty acids, are then boiled with the proper proportion of lye. During the reaction the glycerine of the fat is set free and soap, a compound of the fatty acid with the lye is formed.

If too much alkali is used in making a soap, the soap will contain free alkali which is bad. If not enough alkali is used, the soap will be too greasy. It is up to the chemist in the soap factory to determine exactly the suitable proportions for every sample of soap. In making fine soaps, such as Castile soap and the high grade toilet soaps, olive oil is used instead of common fats and in the manufacture of many soaps, perfume and coloring matters are mixed with the other materials. Ordinary white soaps are made from cotton seed oil. Cheap laundry soaps are made from impure fats obtained from kitchen grease, bones, etc. Glycerine soaps are made by melting hard soap and adding an equal amount of glycerine.

The difference between hard and soft soaps is due to the materials contained in them. Soaps made from caustic potash, or potassium soaps, are soft, while those made from caustic soda, or sodium soaps, are hard.

EXPERIMENT 173—Making Soap

Equipment: Sodium carbonate, calcium oxide, test tubes, measure, salt.

Operation: Make a solution of caustic soda or lye by putting 3 measures of sodium carbonate and 4 measures of calcium oxide into a test tube half full of water and boil for two or three minutes. Allow the tube to cool and when the liquid has settled pour the clear liquid into another test tube. Now add a piece of lard or butter about the size of a marble and boil the liquid again for a few minutes, being careful that the liquid does not jump out of the test tube. You can prevent this by shaking the tube in the flame while heating.

Result: Notice that the lard or butter dissolves very readily in the hot alkali. Now add three measures of common salt and heat the mixture again for two or three minutes. Allow the contents of the tube to cool and notice that the soap separates out as the upper layer. The liquid layer below contains glycerine, salt and impurities. Try washing your hands with the soap you have made.

EXPERIMENT 174—Why Soap Cleanses

Operation: Put a little kerosene oil into a test tube one-third full of water and shake the contents of the tube. Notice that the oil has been broken up into minute droplets. Allow the tube to stand and now notice that the oil runs together and rises to the top of the water.

Repeat this experiment, using the white of an egg in place of the water. Notice that this time the oil is broken up in minute droplets as before, but remains in this condition and does not run together and rise to the top. This forms what is called an emulsion, the oil being emulsified.

Repeat this experiment, using a soap solution in place of the white of an egg and convince yourself that the soap forms a true emulsion with the oil. It is due to this fact that oil, grease and dirt are easily removed from the hands when washed with soap.

EXPERIMENT 175—Making Liquid Soap

Equipment: Glycerine (purchase separately), toilet soap, test tube, test tube holder, alcohol lamp or candle.

Operation: Put a piece of toilet soap about the size of a marble in a test tube and add twice the amount of glycerine. Heat the test tube over a flame until the mixture is all melted together. When this mixture is melted together pour the contents of the test tube into a small bottle and fill the bottle with water.

Result: This will give you a liquid soap which can be used on the hands the same as ordinary soap.

EXPERIMENT 176—Chemical Films and Bubbles

Many times you will have the occasion to use films or bubbles in your laboratory. You may desire to fill a soap bubble with hydrogen gas. You may wish to use films of soap for defraction experiments. The common soap bubble will not stand up well, so below we are giving a formula of a soap that will give a good bubble and that will stand up well.

Pure Castile Soap .... 1 oz.
(Palm-Oil soap is also good)
Distilled Water .... 8 oz.
Pure Glycerine .... 4 oz.

Operation: Cut the soap into thin shavings. Dissolve in the water. When the solution is complete as possible, add the glycerine and stir well. On standing, the mixture becomes clear at the bottom. The clear portion is removed and used.

SOLVING THE HARD WATER PROBLEM IN THE KITCHEN, LAUNDRY, AND BATHROOM

EXPERIMENT 177—Action of Ordinary Soap on Hard Water

Place some vary hard water in a basin and try to make a good washing solution by sudsing with some of your mother's kitchen soap. Note that a large portion of your dissolved soap is used up by the hard water and is turned into insoluble calcium and magnesium soap. These insoluble metallic soaps are sticky and heavy and settle on clothes, your face and hands, in your hair, and on the surface of your wash bowl and bathtub. They cannot be removed by rinsing.

EXPERIMENT 178—Action of Ordinary Soap on Sea Water

Secure a bottle of salt water at the sea shore and shake with some ordinary soap. Note the result.

GUMS, ADHESIVES AND GLUES

Gum arabic, a natural occurring substance, is the essential constituent in the manufacture of gums and adhesive for envelopes and stamps. However, a very good substitute known as dextrin or British gum is now made in large quantities both in this country and in Europe by heating starch and subjecting it to the proper treatment.

Glues are manufactured by boiling with water properly prepared animal matter, such as skins, bones and fish stock, and drying the solution.

A very good household cement for mending crockery, glass, etc., can be made by using water glass or sodium silicate as the binding agent.

EXPERIMENT 179—How to Cement Broken Glass

Operations: By means of a small camel's hair brush, paint the broken surfaces of the glass with water glass. After the water glass starts to harden, which will require a few minutes, press the two broken surfaces together and allow the glass to remain this way for a day or two. If on pressing the two surfaces together a little water glass squeezes out, rub this off with a damp cloth.

Result: Notice after a day or two that the pieces are held very firmly together. Glass mended with water glass will not hold together in contact with water, since water glass is soluble in water.

EXPERIMENT 180—How to Make Envelope or Postage Stamp Mucilage

Equipment: Gum arabic, sugar, measure, alcohol lamp or candle, and starch.

Operation: Put 5 measures of sugar, 2 measures of gum arabic and 2 measures of starch in a test tube one-half full of water. Shake the contents of the tube and allow the tube to stand for five or six hours. Then heat the contents of the tube to boiling and allow the tube to cool. With a soft brush paint some of this mucilage on a piece of paper and allow it to dry. Now when moistened with water you will be able to stick the paper to any surface.

Result: Postage stamp and envelope mucilage is made similar to this. In order to keep this mucilage from going bad, a drop or two of some preservative, such as oil of sassafras or wintergreen, must be added.

EXPERIMENT 181—Labels for Your Bottles

Cut small pieces of adhesive tape from the roll, and stick on the bottles. Then write the name of the chemicals or the formula on the tape with a pen and ink and varnish the label. However, you can purchase gummed labels at your local drug store quite cheaply. LABEL ALL POISON BOTTLES WITH THE WORD "POISON!".

EXPERIMENT 182—Glycerine-Litharge Waterproof Cement

Operation: Litharge (lead monoxide) which is obtainable in any drug store, is mixed with glycerine in various proportions depending on the use to be made of the cement. If a slow setting cement is desired use pure glycerine. If a quick setting cement is desired, dilute the glycerine equally with water.

Result: This cement is useful for making water-tight connections between plumbing fixtures and porcelain. It is very hard when it sets.

INK

EXPERIMENT 183—How to Make Printer's Ink

Equipment: Sodium silicate, powdered charcoal, test tube, stirring rod.

Operation: Put a spoonful of water glass (sodium silicate) in a test tube and add two measures of powdered charcoal, or better, lampblack. By means of the stirring rod, stir the mixture until it is quite uniform. Fill test tube one-third full of water and shake thoroughly.

Result: Printer's ink is made similar to this. Writing done with this ink will last for a long time and is very difficult to remove. This is because the charcoal or lampblack gets into the paper pores where it is held fast.

EXPERIMENT 184—Finger Print Ink and Records

Equipment: Oil paint, lampblack, test tube, sheet of glass, stirring rod.

Operation: Place a few drops of thin oil paint, any color, in a test tube. Add lampblack, a little at a time, while stirring to a thick smooth paste.

Place a small portion of this black paste on a sheet of glass. Spread it around so as to make a very thin, even film on the glass.

Result: Press the fingers one at a time on the plate; make the prints on smooth white paper, being careful not to smear. An album can be made by taking the finger prints of all your friends and marking them.

FERMENTATION

Fermentation is the decomposition or breaking down of complex organic compounds into simpler substances by certain living organisms called ferments. There are many different kinds of organisms or ferments. The most familiar illustration of fermentation is the fermentation of fruits or fruit juices, the decay of vegetable matter or nitrogenous matter, the souring of cider or of milk and that taking place in bread making by action of yeast. Most organic substances, especially the complex substances, are subject to fermentation changes. The chemical changes brought about by ferments consist in the breaking down of complex molecules to form simple groups of atoms. In some cases it consists of rearrangement of the atoms in the molecule to form a compound with different properties. There are many kinds of living organisms or ferments each of which produces its special change. Most ferments are microscopic plants of simple structure which multiply very rapidly when placed in the proper medium or substance. They are variously known as microbes, germs and bacteria. Yeast which is used so commonly in bread making is a microscopic plant found in the air and about certain fruit trees. The yeast cakes which you buy in the store are these same living organisms which have been grown in some suitable culture medium as the potato or corn meal. The above class of ferments are known as organized ferments. The other class of ferments are called chemical or unorganized ferments and are known as enzymes. Examples of these are pepsin, trypsin and diastase.

The action of ferments on organic substances always works best at definite temperatures. Above these temperatures the ferments or organisms are destroyed or their action delayed.

EXPERIMENT 185—Changing Sugar Into Carbon Dioxide by Fermentation

Equipment: Corn syrup, cake of yeast, tumbler.

Operation: Put two spoonfuls of corn syrup or molasses in a tumbler half full of water and stir it up thoroughly. Then dissolve one-quarter of a cake of yeast in a small tea cup full of luke warm water. This is best done by allowing the yeast to stand in the water for about a half hour and then stirring. Now pour the solution of yeast into the tumbler containing the syrup and set the tumbler aside in a warm place for several hours. Notice that after some time a reaction is taking place, small bubbles of gas being given off.

Result: Hold a lighted match over the surface of the liquid in the tumbler and notice that it goes out, due to the carbon dioxide gas which is given off. The syrup or molasses contains a large amount of sugar and during the fermentation produced by the yeast this sugar is changed into carbon dioxide gas.

EXPERIMENT 186—Changing Starch Into Sugar and Carbon Dioxide by Fermentation

Equipment: Starch, cake of yeast, tea cup.

Operation: Put a half teaspoonful of starch in a tea cup and add a spoonful of water. Stir the starch in the water until it is in the form of a paste and then fill the teacup three-quarters full of boiling water and stir for a few moments. Now dissolve one-quarter of a cake of yeast in a teacup one-quarter full of luke warm water and put this solution into the tea cup containing the starch. Set the tea cup away in a warm place and examine it carefully from time to time. Notice the formation of bubbles of gas, due to the liberation of carbon dioxide gas. It will require several days before the starch will be converted into sugar and carbon dioxide.

Result: This illustrates the way starchy foods are converted in the body into sugars which, in turn, are burned up, through the process of assimilation, into carbon dioxide and water.

EXPERIMENT 187—Fermentation Without the Addition of Yeast

Equipment: Glass tumbler, fruit, such as grapes, apple or peach, sugar and water.

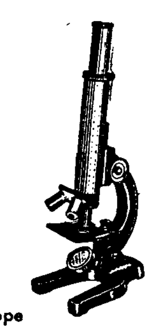
Operation: Crush a few grapes or a portion of other fruit in a glass, add a quarter glass of water and two teaspoonfuls of sugar. Place in a rather warm location.

Result: After a few hours you will notice that small bubbles are forming, rising to the top of the solution and breaking. This shows that fermentation is going on. As no yeast was added to the fruit and water, your natural question would be, what started the fermentation, where did the yeast come from? The answer is that there are little bacteria present in the air which always cling to the surfaces of fruit, etc. They cause the growth of the ferment which started the activity in your fruit and water mixture.

EXPERIMENT 188—Growth of Mould

Operation: Place one-half teaspoonful of ordinary starch in a small pan and add to this a small amount of water. Stir to form a perfect paste and then add about one-third of a cup of boiling water. Set the pan on a stove and continue to boil until a perfect paste is formed. This paste is now bacteria free.

Result: Finally pour the paste into an ordinary tumbler and set it away in a warm place. It will soon become infected with bacteria from the air. Examine it from time to time and notice when mould begins to form. After some mould has formed examine its appearance under a microscope, or magnifying glass.



Gilbert 3-power Microscope

It has been found the ultra-violet rays prevent the growth of mould. These rays are used in the curing of meat.

EXPERIMENT 189—Making Vinegar

Operation: Take about one cup of freshly made sweet cider and add to this two or three drops of strong vinegar. Place the mixture in an ordinary tumbler and set aside, after covering with a plate or piece of cardboard to keep the dust out. At the end of each day taste a little of the cider. You will notice that it becomes more sour daily, and that finally it will have become unfit for drinking, and has turned into vinegar.

Explanation: This is a well-known case of fermentation brought about by means of bacterial organism. The sweetness of the original cider is due to the presence of sugar. By the prolonged action of a special low form of organic life the sugar is transformed into acetic acid. Test your original cider and the final fermented product with a blue litmus paper.

EXPERIMENT 190—Souring of Milk

Operation: Take a tumbler of sweet milk and set it in a warm place where you can watch it. Observe the changes that take place day by day. The milk will become infected with bacteria and undergo fermentation with formation of lactic acid.

Result: This organic acid is formed from the milk sugar in the fresh milk. The lactic acid causes the milk to curdle.

ESSENTIAL OILS AND PERFUMES

The use of perfumes dates back as early as 2000 years ago. For instance, in ancient India the sacred fires in the temples were perfumed with *kusa* or *kus*, the fragrant root of a native grass. This odor is a constituent of many of the present popular perfumes as in the azurea bouquet perfumes. In the early days of China and Egypt perfumes were used in the form of incense and even today the Chinese still burn large quantities of this material as a perfume. All the ancient perfumery materials consisted of crude drugs, flowers, herbs, aromatic gums, resins and woods of the Orient, also many of the fragrant flowers which are used in making some perfumes today. Perfumes are extracted from the crude materials, such as flowers, herbs, etc., in three different ways. First, by treating the finely ground material with oil or fats in which the perfume is soluble. Second, by steam distilling the crude materials during which process the perfume passes over with the steam. Third, by extracting or treating the finely ground materials with volatile solvents such as petroleum, ether, carbon tetrachloride or chloroform. The last process of treating the crude materials with a volatile solvent is the one most commonly used now.

For many years chemists throughout the world have been devoting much research to the synthesis or building up of these natural occurring perfumes until today a new industry has been made possible, the industry of synthetic perfumes and flavoring materials. These materials can be made commercially at a much lower cost than when they are extracted from the natural occurring substances, besides this a great many chemical substances have been found which have most delightful odors and flavors and which at the present time do not exist in nature as such. The synthetic materials of today are successfully utilized in goods of the highest grade and many odor and flavor effects would be impossible without them.

Let us try making some sweet smelling oils and perfumes from flowers, fruits, and spices.

EXPERIMENT 191—Making Rose Water

Operation: Fill an iron or tin vessel half full of rose petals and add enough water to cover the petals. Now place the vessel on the stove and, when the water starts to boil, cover the vessel with a piece of absorbent cloth. When the cloth is wet with water from the condensed steam, remove it and squeeze out the water into a cup. Repeat this operation five or six times and then discontinue the boiling. Smell the water in the cup and notice that it has the sweet odor of the rose.

Result: The rose oil which was originally in the petals volatilized, that is, passed off with the steam and condensed in the cloth along with the water. This is the way oil of rose was made in very early times. Today rose water is obtained by steam distillation in a little different manner.

EXPERIMENT 192—Making Geranium Water

Operation: Prepare geranium water the same way as you prepared rose water in the preceding experiment, using the leaves and flowers of the geranium plant in place of the rose petals. The leaves of the "Rose geranium" make a valuable and lasting perfume. Notice this time that you have a liquid which has the characteristic geranium odor.

EXPERIMENT 193—Making Lilac Water

Operation: Repeat Experiment 191, using lilac flowers in place of rose petals. Notice this time that the liquid obtained has the characteristic lilac odor.

EXPERIMENT 194—Making Violet Water

Operation: You can make violet water the same way as you made rose water in Experiment 191 by using violet petals in place of the rose petals. Notice that the liquid obtained has the characteristic violet odor.

EXPERIMENT 195—How to Make Sachet Powder

Operation: Mix together in a mortar one teaspoonful of cloves, one teaspoonful of cinnamon, one teaspoonful of allspice, and half a teaspoonful of vanilla extract.

Now obtain a cupful of dry geranium or sage leaves and grind them up into a powder by means of the pestle.

Mix the ground geranium or sage leaves with the mixture of cloves, cinnamon, allspice and vanilla extract and place the powder thus formed in a jar with a cover.

Result: Notice the fragrant odor produced from this powder. You can vary the odor of this powder somewhat by using the dried petals of different flowers.

EXPERIMENT 196—How to Make Incense

Operation: Mix together 4 measures of cinnamon, 3 measures of allspice and 5 measures of cloves.

Now pour this mixture on a sheet of paper and mix it with 8 measures of potassium nitrate. Do not grind this mixture.

Result: Heat a little of this mixture in a spoon over a flame and notice the fragrant odor given off which resembles that of burning incense.

EXPERIMENT 197—To Make Wintergreen Extract

Operation: Place some wintergreen berries and tender wintergreen leaves in a cup and grind them up using a spoon. Pour the oil obtained into a test tube and notice the fragrant odor. Taste a little of the oil and satisfy yourself that it is oil of wintergreen.

EXPERIMENT 198—How to Make Lemon Extract

Operation: Remove the skin from a lemon by means of a sharp knife and grind some of it up in a cup the same way as you did in the preceding experiment. Pour the oil into a test tube and notice the fragrant odor. Taste a little of the oil.

EXPERIMENT 199—How to Make Orange Extract

Operation: Remove the skin from an orange, using a sharp knife, and grind some of it up the same as in Experiment 198. Notice the fragrant odor. Taste a little of this oil.

EXPERIMENT 200—Potpourri Sachet — Rose Jars

Operation: Mix together in a cup the following spices or any of them which you can obtain: Half a teaspoonful of cloves, half a teaspoonful of allspice, half a teaspoonful of cinnamon, 2 teaspoonfuls of salt and about half a teaspoonful of extract of vanilla. Grind these materials together until thoroughly mixed.

Now obtain about a quart of dried rose petals and mix this spice which you have just prepared with the rose petals. When it is thoroughly mixed you will find that it gives a very sweet, pleasant odor, and it can be used in sachets or put in jars, thus making rose jars. Whenever the cover is removed from these jars the sweet odor will fill the room.

EXPERIMENT 201—Sage

One of our common plants which has a very sweet odor is the sage. The leaves of this plant when dried retain their odor for a long time, and they can be used to good advantage in sachets and other places where a sweet smell is desired.

PART XII

AGRICULTURAL SCIENCE AND TEXTILES

AGRICULTURE

AS our knowledge of the chemistry of the processes of living organisms—animals and plants—increases, it becomes increasingly difficult to define the boundaries between the animate (living organisms) and inanimate (inorganic or non-living). The significance of this statement will become apparent to any intelligent boy and girl if they will only stop to consider the dependence of our civilization on the successful practice of agriculture. The primitive source of the food of man is the soil, and consequently the successful practice of agricultural science is necessary if man is to secure food for his existence. The farmer occupies, therefore, a most important place in human economy. The energy which is necessary to promote the different chemical changes involved in animal and plant growth is derived from the sun. Deprived of this source of energy, plants would not be able to grow, agriculture would be destroyed, and man would die.

THE CHEMISTRY OF FERTILIZERS

Did you know that chemistry plays such an important part in farming and in the fertilization of soil? No doubt you have often seen the farmer cover his soil in the spring with manure or other forms of fertilizers and later work them into his soil. The reason why these fertilizers were put into the soil is that they make crops grow better. But why is this so? By repeated experiments it has been found that there are certain substances which are very essential to plant growth. We have already learned how plant compounds, such as starch, are built up in the plant by means of carbon dioxide and moisture in the air, with the aid of the sun's rays. On the other hand there are other important substances which are taken up by the plant in the soil to form complex compounds. These are potassium, nitrogen and phosphoric acid. These substances must be introduced into the soil in the form of their soluble compounds as they are taken up by the plant roots by means of absorption.

Nature plays a very important part in the formation of nitrogenous substances in the soil. You probably do not know that in most fertile soils there exists millions of bacteria which have the property of converting decayed nitrogenous organic matter into soluble nitrates which are taken up by the plants. Now it is very important in farming to see that the soil is kept in a condition whereby these bacteria are able to thrive in order to keep the soil fertile. Soils which contain large amounts of decayed organic matter are kept in a moist condition so that the air does not get at them, and are quite often apt to be acidic. The nitrate forming bacteria are not able to live in soils which are acid so that it is very important to see that these acids are destroyed. Acid soils also render other fertilizing substances, such as phosphates, insoluble or in such a condition that the plant is unable to absorb them in the ordinary process of plant growth. This acidic condition of the soil is generally destroyed by neutralization of the acids with lime.

In the ordinary process of farming, since the plants are continually removing the nitrogen, potassium and phosphorus compounds from the soil, it is necessary to replenish these substances with fertilizing material. Now the amounts and kinds of fertilizer to be put into a certain grade of soil depend upon the condition of the soil and kinds of crops to be raised. One class of plants will require more phosphorus than another class. Many formulas have been worked out including the kinds and amounts of fertilizing substances to be used for different kinds of plants.

Kinds of Fertilizers

Soils are fertilized in four ways, namely, by decayed vegetable matter, by rotation of crops, by animal manure and by artificial or commercial fertilizers.

Decayed vegetable matter such as leaves, grass, etc., is often worked into the soil as a natural fertilizer but since they contain only small amounts of the essential fertilizing substances, their use alone would be inadequate.

The second method of fertilization, namely the rotation of crops is widely used in farming and consists in planting fertile land one season with corn, wheat or other crops and the following season with clover, cowpeas or alfalfa. These latter plants have the property of converting nitrogen from the air into soluble nitrates which are restored to the soil to take the place of that used up by the corn, wheat and other crops in the preceding season. The conversion of the nitrogen of the air into nitrates is accomplished by certain germs or bacteria which occur in the tubercles or swellings on the roots of clover, cowpeas, and alfalfa plants.

Natural or animal manures are used exclusively as fertilizing materials as they contain large amounts of nitrogen compounds. They also tend to keep the soils loose so that the air can penetrate them. Guano, cow dung and sheep dung are important natural manures containing relatively large amounts of nitrogen. Guano is the excrement of sea-birds and is found off the coast of Peru. Guano is rich in phosphorus, nitrogen and potash.

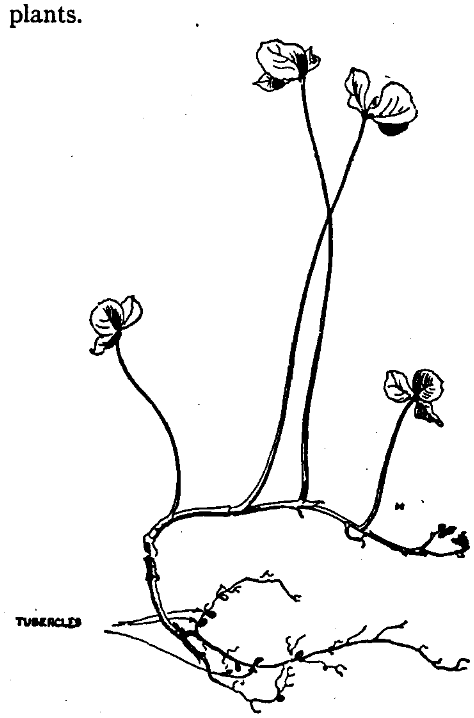
Artificial fertilizers are used very widely today and several important industries are involved in their manufacture. These are chemical compounds or substances that are rich in nitrogen, phosphorus and potassium. Phosphorus is used in these fertilizers in the form of its soluble compound such as calcium acid phosphate. Potassium is used principally in the form of potassium chloride, potassium sulphate and potassium carbonate, while nitrogen is used in the form of ammonium nitrate and sodium nitrate. The commercial fertilizers which are put out on the market are complete fertilizers, such as dissolved potash and phosphates, wood ashes, ground bones, dissolved bones tankage, dry ground fish, nitrate of soda, dried blood, cottonseed meal, linseed meal and castor pomace.

Chemical Plant Foods

One of the important problems of the science of agriculture is to find out which of the elements that occur in plants and animals are necessary for their growth. Hitherto the farmer has been taught that only ten chemical elements are necessary for the growth and normal support of his crops. It has, however, been a well-known fact among agricultural chemists that a far greater number of chemical elements than ten occur in small amounts in fertile soils, and also in the ashes of plants that have been grown under natural conditions. The elements which have been regarded as the only ones essential for the growth of plants are the following: carbon, hydrogen, oxygen, nitrogen, phosphorus, potassium, calcium, magnesium, sulphur, and iron. It was perfectly natural that in the early days of agriculture these elements which occurred in the largest amounts would receive the first attention by agricultural chemists. However, with the development of new and modern methods of chemistry and with greater refinement in methods of research, it was not unreasonable to expect that a very large number of the chemical elements which have been regarded as non-essential would finally prove to be very important factors in plant growth. This has really proven to be the case. It is now known that the element, manganese, for example, is a very essential plant food. It has been shown that plants will assimilate manganese compounds from the soil. It now also seems probable that manganese actually exists in small quantities in all living organisms, and has important functions that have hitherto been unrecognized. All the facts to date show that manganese is an important and necessary factor in the synthesis of the green chlorophyll of plants, and experiments have already been carried out which show that no other one of the common elements — iron, copper, zinc, boron, or arsenic — will replace manganese. Heretofore, it has been believed that iron and magnesium were chiefly concerned in the synthesis of plant chlorophyll. It has now been shown by recent workers in agricultural science that manganese plays a role of importance equal to that of iron in the synthesis of chlorophyll, and that both of these elements are indispensable for the formation of chlorophyll in plants.

EXPERIMENT 202—Nitrogen Forming Bacteria

Pull up a clover of alfalfa plant and examine the roots. Notice that along the roots there are numerous small swelling (see below). These swellings are called *tubercles* and if these were crushed and examined under a powerful microscope you would find millions of small living bacteria in them. These are the bacteria which convert the nitrogen from the air into soluble nitrates which are furnished to the soil and later taken up by other plants.

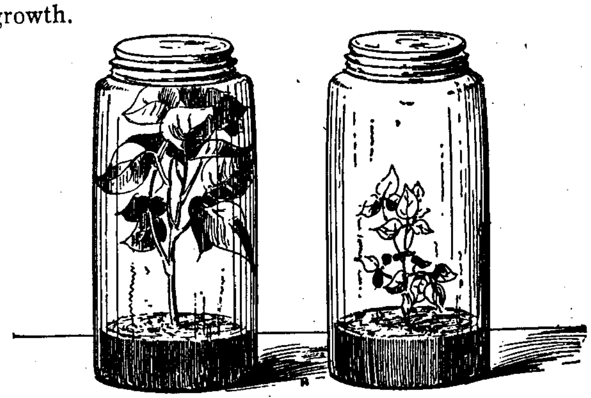


EXPERIMENT 203—The Effect of Carbon Dioxide on Plant Life

Equipment: Sodium bicarbonate, tartaric acid, Mason quart jars.

Operation: Procure two Mason quart jars and place about two inches of damp earth in each. Now plant two or three peas or beans in each jar exactly under the same conditions. Allow the jars to stand until the beans begin to sprout. Then fill one jar with carbon dioxide gas. The carbon dioxide gas is made by putting two spoonfuls of baking soda or sodium bicarbonate in a third empty quart jar and adding a little vinegar or a solution of tartaric acid. After the reaction has gone on for several moments pour the gas from this jar into one of the other two jars, being careful not to allow any of the liquid to get into this jar (see below).

Result: Now seal the two jars containing the sprouted beans or peas and allow them to stand for several days in the sunlight. Label the one containing the carbon dioxide. Examine the jars from time to time and notice that the plants in the jar containing the carbon dioxide grow faster than those in the other jar. This proves that carbon dioxide is essential to plant growth.



AMERICAN FORESTS TO PLAY A NEW ROLE IN ECONOMICS

Forests are growable and replacable natural resources. Consequently wood will always find application in nature's economy and can be called upon more and more to substitute for exhaustible materials. The development of increased consumption of wood and its waste products will benefit our country's economic system. It is believed by our technical men that it can provide an industry of $750,000,000, and furnish employment for 150,000 wage earners. At the present time one-third of the United States is forest land.

Many products that can be manufactured from wood and wood waste are already known, but with the combined efforts of our chemists and chemical engineers, new developments are bound to take place in the future. Some of the products that may be derived from wood waste are rayon, sugar, synthetic rubber, acids, alcohol, dyes, baking powder, tooth-paste, activated charcoal, pulp, paper, explosives, lacquer, paints and varnishes, alkaloids used in medicine, essential oils used in perfumes and chewing gum, plastics, films, glues, yeast, and ink.

TEXTILE CHEMISTRY

Textile fibers may be grouped in two classes according to their source: Vegetable fibers and animal fibers. Cotton, linen, and rayon are all of vegetable origin, the first two being the natural fibers obtained from the cotton and flax plants, respectively, while rayon is a very interesting textile material produced by first dissolving the cotton substance, called cellulose, and then from the viscous liquid, spinning a thin filament which is hardened chemically into a thread so fine and lustrous that it was originally called artificial silk. But since its chemical and physical properties were in many ways unlike true silk, it was soon given a name of its own, rayon. Cellulose is a common material in the woody structure of all plants and trees, so it is not necessary to destroy good cotton fiber to make rayon. It is of further interest that the solubilized cellulose enters the composition of many of our lacquers, and if formed into a sheet instead of fine thread it becomes cellophane.

Silk and wool are produced by animals, the first being the cocoon of the silk worm, and the other, as you all know, is the warm coat of a sheep. These fibers are therefore very different from the cellulose group, and are spoken of chemically as proteins. The chemical differences between these classes of textile fibers are important to remember, especially, in connection with dyeing and with spot and stain removal.

The cellulose of vegetable fibers is very easily damaged by strong acids, but is uninjured by quite severe treatment with alkali. These fibers do not combine with dyes very readily, and it is frequently necessary to deposit the dye on the fiber in a water-soluble form, then convert it to an insoluble product, in order to hold it permanently on the fiber.

Animal fibers on the contrary, are much more resistant to acid and more easily damaged by alkali than the vegetable fibers. They combine with many dyes so firmly that the problem of dyeing colors fast for washing is greatly simplified.

Years ago most dyeing was done with the colored extracts of various plants. There were a limited number of these and many of the bright colors so common today could not be obtained.

Gradually there were discovered ways of making new dyes in the laboratory, building them up from simpler compounds obtained, in many cases, from coal tar. These dyes, sometimes called aniline dyes, now supply every hue of the rainbow and have almost completely displaced the natural dyes.

TESTING TEXTILE FIBERS

EXPERIMENT 204—The Burning Test

Equipment: Alcohol lamp or candle.

Operation: Obtain samples of several fabrics. Take a few threads from a sample and light one from the flame of your alcohol lamp or candle. If the sample is an animal fiber it will not burn rapidly but will char into little black knobs and will produce an odor like burning feathers or hair. If the threads burn readily, leaving a clean white ash, they are vegetable fibers. This test is particularly useful in recognizing whether a sample is silk or rayon.

Explanation: One type of rayon, cellulose acetate sold under the trade name "Celanese," does not burn readily like the other rayons, but more like silk, except that it does not give an odor like burning feathers. Hence an additional test is needed if cellulose acetate is suspected of being present.

EXPERIMENT 205—Acetone Test for Cellulose Acetate

Moisten the textile fibers with acetone. Ordinary animal and vegetable fibers are not changed, but cellulose acetate softens and may even dissolve completely.

EXPERIMENT 206—Microscopic Examination

With experience one can soon learn to recognize most of the fibers by appearance and feel. A magnifying glass or microscope is a great help. With it, wool is seen to have a rough scaly surface while silk is quite smooth. Rayon appears much like silk, but of course the burning test will show them different. The cotton fiber is seen to be an irregular, flattened tube.

EXPERIMENT 207—Wetting Test to Identify Linen

Equipment: Sodium bisulphate, glycerine, test tube, alcohol lamp or candle, test tube holder.

Operation: Linen is more readily wet through than the other fibers. This property is made use of in the following test. Put a small piece of cloth in a test tube half full of water and add one measure of sodium bisulphate. Heat the solution to boiling. (Be sure that the sample is very small so that it doesn't prevent free circulation of the water). Take out the sample, wash it several times with water and dry it. After it is dry, loosen several of the fibers with a pin or needle and moisten them with a drop or two of glycerine. Press the cloth between two blotting papers and then examine very carefully the fibers which are moistened with glycerine.

Result: if they are linen they will be semi-transparent, while other fibers will appear opaque.

EXPERIMENT 208—A Quantitative Test for Fabric Composition

Equipment: Calcium oxide, sodium carbonate, test tube, alcohol lamp or candle, measure.

Operation: Put 6 measures of calcium oxide and 6 measures of sodium carbonate in a tin cup containing two test tubes of water. Boil for several minutes, then let stand to settle and pour the clear liquid, which is sodium hydroxide, into a clean cup while you wash the solid residue from the tin cup. Return the sodium hydroxide solution to a clean cup.

Application: Now take a sample of cloth which you suspect is made of a mixture of animal and vegetable fibers such as wool and cotton. Weigh it carefully, then immerse it in the sodium hydroxide solution and boil it for a few minutes. This should dissolve any wool or silk. Remove the remainder of the sample, rinse it in water, dry and weigh. The loss in weight represents the wool or silk which was in the fabric.

DYEING

You will find it very instructive in the following dyeing experiment if you prepare small patches of cloth containing threads of several kinds. Cut squares of white cotton, about two inches each way. Cotton from an old sheet or pillow case is better than new cloth for by many washings the fibers have become easier to dye. Next obtain coarse white threads or yarns of wool, silk, rayon, and linen. With a needle, stitch one of each of these threads across each square of cotton, using long stitches to keep most of the thread showing on one side. Trim the threads even with the edges of the patch.

EXPERIMENT 209—Dyeing with an Aniline Dye

Aniline dyes are now supplied ready for household use, and the beginner will find these very interesting to experiment with. You will undoubtedly find some on hand in the home laundry.

Dye one of your patches of cloth in one of these dyes, following the directions on the package. Are all the kinds of textile fiber in the patch dyed alike?

EXPERIMENT 210—How to Make Blue Horse Chestnut Dye

Equipment: Ammonia, test tube, alcohol lamp or candle.

Operation: Put several chips of the bark from a chestnut tree into a test tube half full of water and boil four or five minutes. Now add a little household ammonia and boil again for two or three minutes.

Result: Notice the blue color which is formed.

EXPERIMENT 211—How to Dye Cotton Iron Buff

Equipment: Ferric ammonium sulphate, sodium carbonate, test tubes, measure.

Operation: Make a solution of ferric ammonium sulphate by dissolving two measures of the compound in a test tube one-third full of water. Place in this solution a small piece of cotton cloth to be dyed and shake the contents of the test tube thoroughly. Remove the cloth and allow it to dry.

Now dissolve two measures of sodium carbonate in a test tube one-third full of water and place the cloth in this solution. Shake the contents of the test tube thoroughly and then remove the cloth. Wash the cloth with water and allow it to dry.

Result: The cloth will be dyed an iron buff. This color is produced by the precipitation of iron oxide upon the fiber by the alakline salt, sodium carbonate.

SPOT AND STAIN REMOVAL

Chemical treatment is often required in order to remove spots and stains from fabrics. To do this successfully, you need to know, not only what chemical will remove the stain, but how well the fabric will withstand the chemical action. When the fabric is colored, the problem is still more complicated for the dyes may be attacked by the chemical. The following general rules may save you many mistakes.

1. If you use acid on clean cotton, be sure it is very thoroughly washed out before drying.
2. Do not use alkali on silk or wool. If possible, avoid the use of water and clean with a dry-cleaning solvent such as carbon tetrachloride. Many silks may be washed in soapy water, but you should first be sure that they have been dyed with colors fast to washing.
3. Many spot removers act by bleaching with chlorine. Never use these on silk for chlorine damages silk. Chlorine is also very likely to attack dyestuffs. If silk must be bleached, use a solution of hydrogen peroxide made slightly alkaline with sodium silicate.

EXPERIMENT 212—To Remove Iron Rust

Directions: Obtain a small amount of oxalic acid and dissolve four measures in a test tube half full of water. Rub the spot with some of this solution and notice that the stain is removed. Dilute hydrochloric acid can be used in place of the oxalic acid. Remove the oxalic by washing with water.

EXPERIMENT 213—How to Remove Acid Spot

Equipment: Ammonia.

Operation: If acid is accidentally spilled on the clothing, pour a little ammonia on the spots and rub lightly with a cloth. Wash the spot with water in order to remove the salts that are formed in the reaction. This is a case of neutralization, ammonia neutralizing the acid to form a salt.

EXPERIMENT 214—To Remove Alkali from Clothing

Equipment: Tartaric acid.

Operations: If caustic soda is accidentally spilled on the clothing pour some tartaric acid solution or vinegar on the spot and then wash the spot out with water. This is another example of neutralization.

SOURCE OF WOOL

Wool comes from a sheep's back. So does a lot of dirt and other materials picked up on his travels. He also exudes from the pores of his skin oily products which are generated by glands in his body. Only about one-half of the fleece is actual wool fibre. Fifteen per cent is grease, 15 per cent salts, and 20 per cent clay and dirt. Therein lies the reason why the woolen industry calls on the chemist to advise it how to prepare crude fleece for public use as clothing. It is a long chemical.

EXPERIMENT 215—Wool Scouring

Secure if you can a small quantity of wool fleece before it has been subjected to any scrubbing treatment. Examine it and try to clean it by scrubbing with different detergents.

EXPERIMENT 216—Wool Grease

Instead of scouring your fleece with detergents to remove dirt, soak it for several hours with some petroleum naphtha or ordinary gasoline. This will remove the valuable grease known in the drug trade as "lanolin." This wool fat is an important constituent of cosmetic preparations. Finally apply treatment with detergents to remove dirt. These two experiments will serve to reveal to you some of the difficulties facing the woolen manufacturer in preparing his product for public use.

EXPERIMENT 217—Clothing

It takes a sheep and a half to make a man a suit of clothes. Test the fabric used in your shirt and stockings and determine which contains wool.

EXPERIMENT 218—Shoddy

What is the difference between the different forms of woolen cloth—worsteds, serges, flannel and shoddy? Examine a piece of shoddy under your microscope. What is there peculiar about its structure?

SOURCE OF COTTON

Cotton is a plant product and comes from the cotton plant. It is a cleaner commercial product than wool fleece and does not offer as many difficulties in purification. This plant is cultivated in all parts of the temperate and torrid zones, and is the source of enormous industries spread over our globe.

EXPERIMENT 219—Cotton Industries

Try to enumerate thirty commercial products which are manufactured from cotton. Make a chart showing the industries producing these products.

EXPERIMENT 220—Cotton Plant

If possible obtain some raw cotton and examine it. Try planting cotton seeds in your garden. You will be surprised how well they grow even in the northern regions of the United States.

SOURCE OF SILK

Silk is an animal product and is produced by the silk worm known as Bombyx Mori. This worm is cultivated and reared chiefly in China, Japan and Italy. The silk worm feeds on the leaves of the mulberry tree, and stores up a liquid silk solution in his body. After the worm has reached maturity it then exudes this silk solution in the form a fine thread and weaves a cocoon.

EXPERIMENT 221—Cocoons

Secure a silk cocoon and cut it into halves by means of a sharp knife. Observe the dry chrysalis of the original worm encased in the cocoon.

EXPERIMENT 222—Degumming

Take the severed cocoon and boil it for a long time in hot soap solution. Wash with hot water then dry the fiber and note its silky appearance.

CHEMICAL IDENTIFICATION OF TEXTILE FIBERS

In the home will usually be found some form of commercial lyes. These consist essentially of sodium hydroxide and can be used for the following fabric alkali tests. A tablespoonful of lye to a tea cup of water will give a solution of sufficient strength for performing these fabric experiments calling for use of alkali.

EXPERIMENT 223—Identification of Wool

A fiber of wool, when burned, is consumed slowly, a ball-like end being formed. The odor is characteristic of burning animal matter.

EXPERIMENT 224—Identification of Cotton

Burn a fiber of cotton and note that it burns quickly without giving any animal-like odor. The odor resembles that of burning paper.

THE CHEMISTRY OF THE BODY

Man is dependent on the plant kingdom of the universe for his source of food. In this way man utilizes indirectly the energy of the sun. Plants grow up, wither and fall to the ground. They decompose largely into carbon dioxide and water vapor, which are taken up by new plant life. The nitrogen is converted into nitrates by bacteria in the soil and the mineral matter helps to form new plant life.

Sugars are formed in plant life by the condensation of formaldehyde which is produced by the action of carbon dioxide on water with the aid of sunlight. The sugars are then changed into starch by plant ferments or enzymes and these starches are used by animals to support life. In animals these starches are changed back into carbon dioxide and water and the whole process repeats itself. Up to date, man has acquired through research a very extensive knowledge of the chemical changes taking place in this marvelous life cycle.

In animal life we find enzymes similar to those in plant life. Sugar (dextrose) for example, is changed into starch (glycogen) in the liver and this is later turned into dextrose again in the blood by the action of another enzyme. The dextrose in the blood is converted by still further enzymic action into carbon dioxide and water vapor which are carried to the lungs and passed out on breathing. This change takes place in the muscle cells and furnishes the body heat.

How Blood Serves the Body

The blood, which is the most vital part of the body, is composed of water, containing in solution, protein material called fibrinogen, seroglobulin, seralbumin and saline water. Also suspended in this solution are red and white corpuscles. Neutral salts when taken into the body are decomposed, the acid constituent usually hydrochloric acid from salt, going to the gastric juice in the stomach and the alkaline constituent going to the blood. Therefore, the blood is normally alkaline. When blood is exposed to the air it clots. This is due to the fibrinogen in the blood changing to an insoluble compound called fibrin. The blood usually contains about 0.1 per cent of dextrose. The red corpuscles of the blood contain what is known as haemoglobin. This is an enzyme containing iron. It produces some of the very fundamental changes in the body in the same way that chlorophyll does in the plant. This haemoglobin has the property of taking up oxygen from the lungs and later giving up this oxygen during the oxidation of the tissues. During this oxidation heat and energy are formed. The white corpuscles sometimes called leucocytes are really the policemen in the blood for they tend to ward off sickness by destroying the impurities in the blood such as bacteria. When the number of bacteria becomes too great for these white corpuscles we become sick and then we have to resort to medicine to help the action of the white corpuscles.

A very important constituent of the brain and nerve cells is the substance known as lecithin. This is a complex organic compound containing phosphorus. Ordinarily the foods we eat contain small amounts of this substance, thereby furnishing the body with the necessary amount.

You may ask the question, why is the blood normally alkaline? This is easily explained by the fact that it is up to the blood to remove the acidic carbon dioxide gas from the body. In an alkaline condition the blood can easily absorb these waste gaseous products of oxidation and in so doing forms sodium bicarbonate. The sodium bicarbonate in the blood gives up carbon dioxide and water vapor in the lungs. This leaves sodium carbonate which is the alkaline constituent of the blood and the blood is now ready to pass again through the body and take up more carbon dioxide.

The teeth differ from the other parts of the body in that they are more resistant to chemical action, namely, that of acids. The mouth of a normal person is alkaline and in this state the teeth are well protected. Decayed matter if allowed to cling to the teeth is converted into acids which slowly but effectually remove the enamel from the teeth and cause them to decay. The enamel of the teeth is composed chiefly of calcium fluoride. You can readily see, therefore, that it is very important to brush the teeth every day and the use of a mild antiseptic as a mouth wash is very effectual.

EXPERIMENT 225—How to Test for Acid Mouth

Moisten a small piece of blue litmus paper with the tongue and see if the litmus paper turns red. If it does, the mouth is acid. Acid mouth is usually caused by an upset stomach or by decayed teeth. Normally the mouth should be neutral or only slightly alkaline.

Health With Food and Exercise

It is necessary to say at this time a few words in regard to body health. There are several influences which tend to keep up a healthy normal body. These are exercise, fresh air, cheerfulness, cleanliness, sunlight, pure food and pure water. The relationship between chemistry and exercise may be shown if we consider that many chemical reactions take place much more rapidly and with better results if the solution in which the reaction is taking place is heated or stirred continually. Just so in the body. By exercising the body the blood moves faster, the stomach digests our food better, we breathe much more efficiently and as a consequence the waste products of the body are more effectually eliminated and we feel much better.

Sunlight has several effects on the body. It is a powerful germicide and will kill many organisms which are the cause of disease. It also acts as a blood stimulator. You probably have noticed how red the body becomes in the summer time when exposed to the sun's rays.

Importance of Body Cleanliness

Cleanliness is a very important contribution toward health. Dirt is a very good medium or culture for bacteria, so that it is quite necessary to keep the body clean. We have always spoken of the importance of keeping the pores of the skin open so that impurities can come out in the form of perspiration and that the pores may breathe in a certain amount of fresh air. Chemistry has contributed toward this influence on the health of the body by furnishing such substances as borax, soap, soda, ammonia, antiseptics and synthetic remedies.

Pure fresh water is a very important factor in every day life. Natural waters contain small amount of mineral salts, principally sodium chloride. These are not injurious to the health but rather beneficial. Waters, however, are apt sometimes to contain bacteria which may be very injurious to the health. If a water is suspected of containing bacteria it should be boiled. On a commercial scale water is purified by several processes, namely, treating the water with ozone, bleaching powder, chlorine, ultra-violet light or by distillation. Any of these methods are good if used under the right conditions.

Breathe Plenty of Fresh Air

The subject of fresh air is an old one, but one which a great many people disregard. It is very important that the lungs obtain the normal amount of oxygen necessary in the oxidation processes which take place in the body. The effect of poisonous gases when breathed into the lungs was very ably demonstrated during World War I. Impurities in the air such as poisonous gases, etc., have a very remarkable affinity for the blood. They react with the blood to form much more stable compounds than oxygen does. Consequently the blood cannot take up the normal amount of oxygen required in the body, and as a result the blood becomes congested. In the case of extreme poisoning the lungs become so congested that we are unable to breathe further, and death follows. In a poorly ventilated room the carbon dioxide given off from the lungs soon becomes in excess of the oxygen present in the room, with the result that we become drowsy or sleepy. This effect seems to be due to a lowering of the oxygen content in the room and also to a rise in temperature caused by the body heat.

Cheerfulness is not directly related to chemistry but is mentioned here because of the influence it has on the health of the body. A person who is cheerful, particularly during the time of eating, is apt to be less troubled with indigestion than a person who is constantly worrying. Loss of appetite is quite often due to just this thing. Indigestion and loss of appetite are controlled to a large degree by the nerves, which are apt to break down in a person who is not cheerful.

The truth of the old saying that an apple a day keeps the doctor away is now vouched for by science. Apples contain vitamin C, which has been found necessary and valuable in promoting health and protection from diseases like scurvy. All nutrition authorities now report that apples, while they do not contain as much of this vitamin as tomatoes or oranges, do supply an important amount of it, particularly if eaten raw, skin and all.

PENICILLIN

The production of *penicillin* by molds stands out today above all fermentation processes. This so-called "wonder drug" has aroused world wide attention because of its outstanding antagonism to bacterial organisms which have been the cause of countless deaths in the past. It is especially effective in preventing infections caused by wounds, and consequently the recent war brought about a large production of this valuable product. It will be difficult to make any forecast as to the contribution which *penicillin* will make to our national economy in the future.

LIST OF CHEMICALS, THEIR FORMULA, AND PURCHASE PRICE

When ordering supplies, please list number with chemical name. Kindly enclose check, stamps, money order, cash or U.S. Postal notes with order. Add 5% to cost on all orders originating from Zone 2*. We pay postage except on C.O.D. orders.

No.	Name	Price
74	Acetone $(CH_3)_2CO$	$ .10
65	Acetic Acid CH_3COOH	.10
1	Aluminum Sulphate $Al_2(SO_4)_3$	.10
69	Ammonia NH_4OH	.05
2	Ammonium Chloride NH_4Cl	.10
6	Ammonium Nitrate NH_4NO_3	.10
4	Borax $Na_2B_4O_7 10H_2O$	.10
5	Boric Acid H_3BO_3	.10
9	Calcium Carbonate $CaCO_3$	.10
8	Calcium Chloride $CaCl_2 6H_2O$	.10
7	Calcium Hypochlorite CaCl(OCL)	.10
12	Calcium Monophosphate $Ca(H_2PO_4)_2H_2O$	.10
11	Calcium Oxide CaO	.10
13	Calcium Sulphate $CaSO_4\ 2H_2O$	.10
14	Calcium Sulphide Paper	.10
10	Camphor Gum $C_{10}H_{16}O$	.10
15	Carbon Tetrachloride CCl_4	.10
75	Chrome Alum $Cr_2(SO_4)_3\ K_2SO_4 24H_2O$	.10
16	Cobalt Chloride $CoCl_2 6H_2O$	.10
17	Cochineal	.10
18	Congo Red Paper	.05
64	Collodion	.10
19	Copper Strip Cu	.05
20	Copper Sulphate $CuSO_4 5H_2O$	.10
68	Denatured Alcohol C_2H_5OH	.05
22	Ferric Ammonium Sulphate $(NH_4)_2SO_4 Fe_2(SO_4)_3 24H_2O$	.10
21	Ferrous Ammonium Sulphate $(NH_4)_2\ SO_4\ FeSO_4\ 6H_2O$	.10
24	Glycerine $CH_2OHCHOHCH_2OH$	.15
23	Gum Arabic	.10
27	Insulated Copper Wire	.10
6	Litmus Paper	.05
28	Logwood	.10
30	Manganese Dioxide MnO_2	.10
31	Manganese Sulphate $MnSO_4 4H_2O$	.10
29	Magnesium Sulphate $MgSO_4 7H_2O$	.10
26	Nickel Steel Wire	.10
32	Nickel Ammonium Sulphate $(NH_4)_2\ SO_4\ NiSO_4 6H_2O$	.15
59	Nigrosine	.10
33	Phenolphthalein $(C_{20}H_{14}O_4)_2\ COC_6H_4CO$	.20
34	Potassium Nitrate KNO_3	.15
35	Potassium Permanganate $KMnO_4$	.10
37	Powdered Charcoal C	.10
36	Powdered Iron Sulphide FeS	.10
38	Powdered Iron Fe	.10
39	Powdered Magnesium Mg	.15
40	Powdered Zinc Zn	.10
61	Red Saunders	.05
42	Sodium Bicarbonate $NaHCO_3$	.10
43	Sodium Bisulphate $NaHSO_4$	.20
44	Sodium Bisulphite $NaHSO_3$	.15
45	Sodium Carbonate $Na_2CO_3\ 10H_2O$	.10
76	Sodium Chloride NaCl	
46	Sodium Ferrocyanide $Na_4Fe(CN)_6\ 10H_2O$	.10
47	Sodium Iodide Solution NaI	.10
48	Sodium Silicate $Na_2\ SiO_3$	.10
49	Sodium Sulphocyanate NaCNS	.15
50	Sodium Thiosulphate $Na_2S_2O_3\ 5H_2O$	.10
73	Strontium Chloride $SrCl_2\ 6H_2O$	.10
51	Strontium Nitrate $Sr(NO_3)_2\ 4H_2O$	.10
52	Sulphide Test Paper	.10
53	Sulphur S	.10
54	Tannic Acid C14 H10 09	.20
55	Tartaric Acid $COOH(CHOH)_2COOH$	.20
63	Tincture of Benzoin	.10
56	Turmeric Paper	.05
57	Zinc Strip	.15

MINERALS

No.	Name	Price
X1500-O	Apatite	.10
X1500-J	Calcite	.05
X1500-C	Chalcopyrite	.15
X1500-K	Fluorite	.10
X1500-A	Galena	.10
X1500-Q	Garnet	.10
X1500-L	Halite	.05
X1500-E	Magnetite	.10
X1500-H	Malachite	.10
X1500-P	Muscovite	.05
X1500-M	Orthoclase	.10
X1500-D	Pyrite	.10
X1500-F	Pyrolusite	.10
X1500-R	Quartz	.10
X1500-G	Sphalerite	.15
X1500-B	Stibnite	.10
X1500-N	Talc	.10

APPARATUS AND EQUIPMENT

No.	Name	Price
X1494	Alcohol Lamp	.20
P1549	Beaker	.25
P860	Black Cloth for Ink Trick	.05
X5618	Blow Pipe	.25
P1599	Candle	.03
P1583	Carbon Electrodes	.10
P1574	Charcoal Block	.20
P1598-A	Cork with Hole two for	.05
P1522	Filter Paper Disc six for	.05
P1548	Flask	.25
X1584	Gas Generating Bottle (Glass only)	.10
X1494-G	Gas Generating Bottle (complete)	.25
P1578	Glass Funnel	.15
P1503	Glass Rod	.05
P1504	Glass Tubes 4"	.05
P1588	Porcelain Mortar	.10
P4727	Horseshoe Magnet	.10
P1455-A	Metal Test Tube Rack	.10
P1589	Porcelain Pestle	.15
P5051	Quill Brush	.03
P1560	Right Angle Tube — Long	.05
P859	Ring for Ink Trick	.05
P-57-A	Rod	.02
P3308	Rubber Coupling	.01
P3309-A	Rubber Tubing, 2 ft.	.35
P1495	Cork Stopper, 2 Holes	.05
P1496	Cork Stopper, 1 Hole	.05
X1531	Scale — Complete	.50
P1577	Short Right Angle Tube	.05
P1580	Small Shovel	.02
P1452	Spoon	.05
P1502	Test Tube	.05
P1462	Test Tube Brush	.10
P1563	Test Tube Holder	.10
X1547	Thermometer	.15
X861-A	Wand	.05

MANUALS

No.	Name	Price
M512A	Glass Blowing Manual	.25
M1272A	Mineralogy Manual	.25
M447A	Chemical Magic Manual	.25
M1706A	Beginners Chemistry Manual	.25
M2394	Junior Chemistry Book	.40
M2395	Senior Chemistry Book	.60

*"Zone 2 consists of Arizona, New Mexico, California, Washington, Oregon, Idaho, Nevada, Utah, Colorado, Wyoming, Montana and El Paso, Texas."