EXPERIMENT 209—Zinc hydroxide
Place a small piece of zinc and 1 measure of sodium bisulphate in a test tube half full of water and a few drops of sodium hydroxide solution. A very pretty blue precipitate of the hydroxide is formed which very soon changes to red.

EXPERIMENT 210—Cobalt hydroxide
Dissolve 1 measure of cobalt chloride in a test tube half full of water and a few drops of sodium hydroxide solution. A very pretty blue precipitate of the hydroxide is formed which very soon changes to red.

EXPERIMENT 211—Magnesium hydroxide
Dissolve 2 measures of magnesium sulphate in a test tube half full of water and a few drops of sodium hydroxide solution. A heavy white precipitate of magnesium hydroxide is formed.

EXPERIMENT 212—Calcium hydroxide
Dissolve 2 measures of calcium chloride in a test tube half full of water and a few drops of sodium hydroxide solution. A heavy white precipitate of calcium hydroxide will form. This calcium hydroxide is the same substance that is obtained when slaked lime is formed.

EXPERIMENT 213—Chromium hydroxide
Dissolve 2 measures of chrome alum in a test tube 1/4 full of water. To this add a few drops of sodium hydroxide solution. A beautiful green precipitate of chromium hydroxide is formed.

Potash or potassium carbonate is an important compound and is essential for the growth of plants. It is absorbed by plants from the soluble compounds in the earth or soil.

EXPERIMENT 214—Obtaining potash from wood ashes
Potash was formerly obtained by leeching out the soluble material in wood ashes with water. Put 4 spoonfuls of wood ashes in a glass half full of water and stir well for 5 minutes. Now allow this mixture to stand for a few minutes and then filter the liquid into a test tube. Test this liquid by adding 2 or 3 drops of phenolphthalein solution. Notice that the solution turns red. Potassium carbonate is found in wood ashes and is removed in this way. By evaporating the liquid down to dryness, one could obtain the solid compound.

ALKALINE EARTH METALS, CALCIUM, STRONTIUM, BARIUM

The term “alkaline earth” was originally applied to the oxides of these metals because they resemble both the alkalies and the earths, the latter term being used for oxides of aluminium 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 oxygen, 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) are extensively used.}

EXPERIMENT 215—Burning limestone to quick lime
Grind a few pieces of limestone, marble or oyster shells and break them up into smaller pieces, and then burn the powdered limestone in a crucible. The compound remaining is quick lime or calcium oxide. 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.

EXPERIMENT 216—Slaked lime or calcium hydroxide
To a test tube add 1/2 inch of powdered lime and 5 or 6 drops of water. Notice that the mixture begins to boil. Now fill the test tube 3/4 full of water and shake the contents of the tube. Filter off the liquid and test it with red litmus paper. This liquid is a solution of calcium hydroxide. It is a weak base and is used in medicine as a base or liquid, which is used in stone foundations for buildings, is made by mixing together sand and slaked lime. On exposure to the air mortar sets or becomes hard. This substance is used in the building of houses.

EXPERIMENT 217—Making mortar
Take three measures of calcium oxide and mix this together thoroughly with three parts 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 the paste becomes hard and sets.

EXPERIMENT 218—To make Plaster of Paris
Take a spoonful of calcium sulphate and heat over the alcohol or gas flame for ten minutes to a high temperature. Allow to cool and then empty the contents of the crucible on a sheet of paper. This is plaster of Paris.

EXPERIMENT 219—Making a cast with plaster of Paris
Mix up the plaster of Paris which you made in experiment 218 into a fine powder and mix the plaster with water to the consistency of a paste. Place a coin slightly greased with oil on a paper. Spread some of this paste on the coin. Press the paste on the coin to displace any air bubbles and leave the plaster undisturbed until it hardens. Then remove it and you will notice a perfect imprint of the face of the coin on the plaster.

It is important to call attention to the application of magnesium chloride and magnesium sulphate in the production of plaster cement used for stucco construction and in the formation of oxycarbides and oxyphosphates, respectively. These cements form the basis of stucco, synthetic wood, and sound absorbing materials. This material has found extensive use for the construction of office buildings and public buildings.
Calium sulphide when it contains traces of sulphones of some of the other alkaline earths resembles each other in chemical properties, they differ in that they give a flame. Calcium compounds give a red flame, strontium gives a green flame, and barium gives a green flame.

**EXPERIMENT 221—Flame test for metals**

Clean the steel wire and make a loop in the cleaned wire. Heat the loop in a flame until there is no color given to the flame by the wire. Dip the clean wire into some powdered calcium chloride and then heat the flame of the wire. Notice the red flame of calcium.

**EXPERIMENT 222—Red fire**

Mix together thoroughly in a pan one measure of strontium nitrate, two measures of potassium nitrate, one measure of sulphur and two measures of powdered coal. Make into a small pile and keeping the face at a safe distance light it with a match. Notice that the mass takes fire readily and burns with a red light to the strontium. The sulphur and charcoal act as combustible materials, and the potassium nitrate furnishes oxygen for the reaction.

**EXPERIMENT 223—How to make green fire**

Zinc, when in the form of a powder, burns with a green flame.

Repeat experiment 222, using two measures of powdered zinc, two measures of potassium nitrate, two measures of powdered charcoal and one measure of strontium nitrate. Notice that this mixture when ignited will burn with a green flame. These mixtures can be used as a substitute for gunpowder in small fireworks, since the zinc, charcoal and potassium nitrate will give off oxygen for the reaction.

**EXPERIMENT 224—How to make yellow fire**

Sodium, when burned, produces a yellow flame.

Repeat experiment 223, using one measure of dry sodium chloride, two measures of potassium nitrate, one measure of sulphur and two measures of powdered charcoal. Notice that the mixture when ignited will burn with a yellow flame.

**EXPERIMENT 225—Preparing a safe explosive**

Chemicals needed: iodine crystals and ammonium. Apparatus you will need: a beaker, filter paper and paper.

Crush several of the iodine crystals and place them in a clean test tube. Add one teaspoon of ammonium and let this stand for half an hour. The reaction is the same as the before experiment. The iodine is dry just a screen on the filter paper will explode the crystals.

**COLOR IN FIREWORKS**

The beautiful red of Roman candles, sky rockets, and even lighted matches, are produced by the use of red in the compound. The red is the result of a chemical reaction and is not permanent. The red is due to the presence of a compound of strontium and potassium. The compound is called strontium nitrate, and it is a chemical compound that gives a red flame when ignited. The compound is used in the manufacture of red in fireworks. The red is produced by the reaction of the strontium and potassium with oxygen in the air. The red is visible because it is produced in a very concentrated form.

**EXPERIMENT 226—Pharaoh's serpent**

Two parts potassium dichromate, this can be secured in a trip, can be made by mixing one part of potassium nitrate and one part of sugar. The ingredients must be perfectly dry and then mixed. The paper cones can be made and filled with a mixture of potassium nitrate and the cone must be ignited. The paper cone will fly up in the air and then descend in a ball of fire. The paper cone is called a "serpent" and it is a spectacular sight to see.
ALUMINUM, ZINC, MAGNESIUM

These are three very important metals and find wide application in industry. Aluminum is a very important metal because it is very malleable, is easily worked, and is more rigid than the same weight of other metals. It is used for 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 also a conductor in electric power lines.

Aluminum forms valuable alloys with steel or magnesium. The compounds of aluminum are important. The silicates are used extensively in the manufacture of cement, brick, tile, earthenware, pottery and porcelain. Valuable gems contain aluminum. The ruby, sapphire and topaz are traces of aluminum oxide, containing small amounts of certain metallic oxides. These give the gem its color.

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

**EXPERIMENT 230—Colored aluminum lake**

Put two measures of cochineal in a test tube half full of water, boil over. Now add one measure of aluminum sulphate and stir until it dissolves. When the solution is clear, add one measure of sodium carbonate in another test tube half full of water. Stir until dissolved. Now mix the two solutions and notice the formation of a red precipitate. This is aluminum hydroxide and it is red. The process is used in dying and in clarifying wine.

**EXPERIMENT 231—Action of sodium carbonate on aluminum sulphate**

Dissolve one measure of aluminum sulphate in a test tube half full of water. Add one measure of sodium carbonate to another test tube half full of water and stir until dissolved. Now mix the two solutions and notice the formation of a white precipitate. This precipitate is aluminum hydroxide. Aluminum carbonate has been prepared, because it is a salt of a weak acid and a weak base. Cork the tube and let it stand overnight. The gas given off in the reaction is carbon dioxide gas. Alum baking powder or cream of tartar substitute, as it contains sodium bicarbonate, starch and potassium alum. These are used to keep the materials dry.

Zinc is used in batteries and as a coating for other metals to protect the oxygen of the air. Galvanized iron is made by dipping iron into molten zinc, allowing to cool. Zinc forms several important alloys. Zinc also combines readily with certain other elements as oxygen and sulphur.

**EXPERIMENT 232—Action of zinc on alkalis**

To a test tube half full of water and hydroxide, prepared as directed in Exp. 202, add one measure of powdered zinc. Warm, if necessary, to start the reaction. Notice the gas bubbling off. This is hydrogen.

Magnesium occurs abundantly in nature as the carbonate. Metallic magnesium is a very important metal because of its extreme lightness and strength. Magnesium is a rival of aluminum in construction demanding a light metal. Many of the hollow plaster of Paris cones are made of magnesium, which is being extensively used.

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

**EXPERIMENT 233—How to make a white flashlight powder**

Mix together on an old pan one measure of powdered magnesium and one measure of potassium nitrate. Do not rub or grind the mixture. Now place one measure of the mixture on top of the mixture and carefully light the sulphur. The sulphur will burn slowly the mixture will flash, giving off a very brilliant light. In this same experiment, leaving out the sulphur and using a fuse made by soaking a piece of string in a strong solution of potassium nitrate and allowing the string to dry, will produce a very bright light. Set off the mixture with a fuse, place the mixture on one end of the string and light the other end with a match.

**EXPERIMENT 234—How to make a red flashlight powder**

Mix experiment 233, using one measure of strontium nitrate in place of the alum nitrate. Notice the brilliant red light produced when the mixture is ignited.

**EXPERIMENT 235—How to make a green flashlight powder**

Experiment 233, using the following proportions of substances: one measure of potassium nitrate, two measures of powdered magnesium, one measure of boric acid and one measure of sulphur. Notice this time that a bright green flash is produced. It will not allow any of the above flashlight mixtures to remain in the air too long, as the salts take up water from the air and therefore will not ignite quickly.

**EXPERIMENT 236—Making sparklers**

Mix one measure of potassium nitrate, two measures of powdered magnesium, one measure of boric acid and one measure of potassium chlorate. Mix the two powders together on a sheet of paper. Light the nitrate and drop in the mixture of magnesium and boric acid. Do not rub or grind the mixture. Notice how many bright flashes are produced and how the salts take up water from the air and therefore will not ignite quickly.

**EXPERIMENT 237—Thermite fusion**

A well-known experiment on thermite can always be depended upon to arouse interest among students of science. The experiment should not be undertaken, however, without the advice and direction of an older person, and should not be conducted in surroundings where there is danger of fire and injury to property. It is a very safe experiment to perform, but the necessary chemicals are not supplied in the student’s set, and must be purchased in the market. For the experiment, it is first necessary to prepare a plaster of Paris cone. This is made by mixing the inside of a large funnel with vaseline. A large paper cone is placed inside the funnel and it is also coated with vaseline. A very small hole is made in the top of the cone. The cone is placed in the funnel and a small plug of vaseline is placed on the top of the cone. The cone is ignited and the mixture is poured into the cone. The mixture is allowed to cool and the cone is placed in the funnel. When the cone is placed in the funnel, it is very easily slipped from the funnel.
The cone is placed in a large ring on a ringstand, at the base of which is sand. A piece of paper is stuffed into the opening at the bottom of the cone to prevent the ignition mixture from dropping out. The cone is then filled with a mixture of aluminum powder and iron oxide. Other oxides, of course, may be substituted. To ensure a good yield of molten metal, a liberal amount of iron is added to the mixture.

On top of the mixture a very small heap of an oxidizing agent is placed. A piece of magnesium ribbon is then stuck into the oxidizer, the cone is covered with asbestos, containing a small hole for the magnesium strip, and the magnesium ribbon is ignited.

This demonstration is most dazzling and impressive when shown in a dark room.

**SILICATES**

**EXPERIMENT 238—Siliceous acid**

Put ½ inch of water glass in a test tube and add water until the tube is half full. Shake to mix the liquids.

Now add two or three drops of water glass to the solution of nickel ammonium silicate and you will get a beautiful green precipitate of nickel silicate.

**RIMENT 244—Ferrous silicate**

Solve the two measures of ferrous ammonium sulphate in a test tube half full of water. To this solution add two or three drops of water glass and a thick precipitate of silicate will be formed.

**RIMENT 245—Ferric silicate**

Solve the two measures of ferric ammonium sulphate in a test tube half full of water and add two or three drops of water glass. A very pretty reddish-brown precipitate of silicate will be formed.

**RIMENT 246—Tin silicate**

Solve one measure of sodium bisulphate, one measure of ammonium chloride, and a piece of tin metal into a test tube. Add five or six drops of water and heat the mixture by means of a bunsen burner until it boils for two or three minutes. Pour the clear solution into another test tube and add water until the tube is one-quarter full. Add two or three drops of water glass and in a few moments a thick white precipitate of silicate will form.

**RIMENT 247—Chromium silicate**

Solve two measures of chrome alum in a test tube full of water. Add two or three drops of water glass and a beautiful thick green precipitate of chromium silicate will form immediately.

**RIMENT 248—Cobalt silicate**

Solve one measure of cobalt chloride in half a test tube of water and add three drops of water glass. In this case a beautiful blue precipitate of cobalt silicate is formed.

**RIMENT 249—Copper silicate**

Solve two measures of copper sulphate and one measure of copper sulphate in a test tube and fill the tube half full of water. Heat the liquid for a few moments. A clear blue solution of copper sulphate is obtained. Add the solution of copper sulphate to two or three drops of water glass and examine the blue precipitate of copper which is formed.

**RIMENT 250—Manganese silicate**

Solve two measures of manganese sulphate in a test tube half full of water and heat the liquid for a few moments to completely dissolve the solid. Add the solution to two or three drops of water glass. A pale pink precipitate of manganese silicate will be formed.

**RIMENT 251—Magnesium silicate**

Solve two measures of magnesium sulphate in a test tube half full of water. To this solution add two or three drops of water glass. A white precipitate of magnesium silicate will be formed.

**RIMENT 252—Calcium silicate**

Solve three measures of calcium chloride in a test tube half full of water, and shake the solid is dissolved. Now add two or three drops of water glass and you will get a beautiful white precipitate of calcium silicate.

**EXPERIMENT 239—Sodium silicate (water glass)**

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

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

**EXPERIMENT 240—Strontium silicate**

Dissolve two measures of strontium chloride in half a test tube of water and add two or three drops of water glass. A bulky white precipitate will form and upon shaking the test tube the precipitate will fill the whole tube.

**EXPERIMENT 241—Zinc silicate**

Place a small piece of zinc metal and two measures of sodium bisulphate in a test tube. Fill the tube half full of water, heat the solution for a moment and some of the zinc dissolves. Now hold the tube in a glass of cold water for one or two minutes until it becomes cool.

Add to the solution of zinc sulphate thus formed two or three drops of water glass and note the thick white precipitate of aluminum silicate which is formed.

**EXPERIMENT 242—Aluminum silicate**

Place two measures of aluminum sulphate in a test tube and fill the tube half full of water, shake to completely dissolve the solid. Now add two or three drops of water glass and note the thick white precipitate of aluminum silicate which is formed.

**EXPERIMENT 243—Nickel silicate**

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

EXPERIMENT 253—Zinc ferrocyanide
Dissolve two measures of sodium ferrocyanide in a test tube half full of water. In another test tube put a small piece of zinc metal and two measures of potassium bisulphate. Fill the test tube 1/4 full of water and shake until the solids are dissolved. Now add a few drops of the solution of sodium ferrocyanide solution from the first tube and a white precipitate of zinc ferrocyanide will be formed.

EXPERIMENT 254—Aluminium ferrocyanide
Dissolve one measure of aluminium sulphate in a test tube half full of water and add a few drops of the solution of sodium ferrocyanide prepared before and a light blue precipitate of aluminium ferrocyanide will be formed. If this precipitate appears early it is an indication, as it shows that there is a trace of iron in the aluminium sulphate.

EXPERIMENT 255—Nickel ferrocyanide
Dissolve one measure of nickel ammonium sulphate in a test tube half full of water and add a few drops of sodium ferrocyanide prepared before and a green precipitate of nickel ferrocyanide will result.

EXPERIMENT 256—Ferrous ferrocyanide
Dissolve one measure of ferrous ammonium sulphate in a test tube half full of water and add to this solution two or three drops of sodium ferrocyanide solution and the resulting bluish white precipitate which forms is the same thing as Turnbull's blue.

EXPERIMENT 257—Ferric ferrocyanide
Dissolve one measure of ferric ammonium sulphate in a test tube half full of water and add a few drops of sodium ferrocyanide solution. The deep blue precipitate which results is called Prussian blue, which is ferric ferrocyanide.

EXPERIMENT 258—Manganese ferrocyanide
Dissolve one measure of manganese sulphate in a test tube half full of water and add to this solution a few drops of sodium ferrocyanide solution and a colourless precipitate of manganese ferrocyanide is formed.

EXPERIMENT 259—Cobalt ferrocyanide
Dissolve one measure of cobalt chloride in a test tube half full of water and then add a few drops of sodium ferrocyanide solution and a white precipitate of cobalt ferrocyanide is formed.

EXPERIMENT 260—Chromium ferrocyanide
Dissolve one measure of chrome alum in a test tube 1/4 full of water. Then add a few drops of sodium ferrocyanide solution and a deep green color will form, indicating the formation of chromium ferrocyanide. Notice that in this case the substance is not in the form of a precipitate as the chromium ferrocyanide is soluble.

EXPERIMENT 261—Tin ferrocyanide
Put one measure of bismuth sulphate, one measure of ammonium chloride, and a small piece of tin metal in a test tube and add five or six drops of water. Heat the solution, allowing it to boil for two or three minutes to dissolve some of the bismuth. Now pour the clear solution into another test tube and add a few drops of sodium ferrocyanide solution, allowing it to stand for some time. Note the light bluish green precipitate of tin ferrocyanide which is formed.

PART III

Organic Chemistry and Its Commercial Application to the Industries

CARBON

Carbon chemistry is based on our knowledge of the properties and reactions of the element carbon. This element is found in nature in the free condition in several forms. The diamond is practically pure carbon, while ordinary coal graphite contains small percentages of other substances besides carbon as mineral matter. Naturally occurring compounds of carbon are of wide occurrence in nature and are found in the form of gases, liquids or oils and solids. Carbon dioxide is the familiar gaseous compound of carbon. Manufactured illuminating gas, natural gas from wells and petroleum are all composed chiefly of organic compounds of carbon and hydrogen.

Carbon compounds, especially calcium carbonate, constitute a very large proportion of all rocks and some form of mineral carbonate are found in most localities. Building stone—marble—is a very pure form of calcium carbonate. Carbon contains a large percentage of all living organisms, both plant and animal, and is essential to the manufacture of organic products as proteins, fats, sugars and natural oils. All these materials are widely utilized by man as food and for the manufacture of useful commercial products. At the present time more than 300,000 organic compounds are known, and the possibilities of new creations as the science of organic chemistry is developed are unlimited. Of all the elements occurring in nature, carbon is the most commonly associated with life itself. Our present world could not exist without it.

As a modified form of impure carbon, and is used as a fuel in operating steam engines, and also in smelting processes for refining ores. It is made by heating bituminous coal, or as it is more commonly known, soft coal until all volatile products in the form of gas have been expelled. This heating process is conducted without excess of air, or oxygen, and the product is carried out in a commercial scale in large ovens. The volatile or gaseous products thus obtained are converted into the raw materials for the manufacture of illuminating gas, and the remaining char is used as a source of carbon for manufacture of illuminating gas, and in the production of charcoal and coke. A product of wood charcoal obtained from the burning of wood which is subjected to destructive distillation. It is a product of the destructive distillation of carbonaceous materials, such as wood and sugar, and is a very important source of raw materials for the manufacture of illuminating gas. The quality of wood charcoal obtained depends on the kind of wood which is subjected to destructive distillation. It is very efficient form of absorbent carbon, meeting the exacting requirements of the trade. If air was admitted during the distillation process, the charcoal and gas products would be burned up completely with formation of carbon dioxide gas.

Another form of carbon is graphite. This is the black substance which forms the lead pencils. It is sometimes referred to as lead, but this is not correct. It is a form of carbon, but a modified form of carbon, and can be made by subjecting carbon to intense heat. At about 4000° C. carbon vaporizes, and this vapor on condensing upon a surface 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 ink and pencils.

Lamp black, an amorphous form of carbon, finds wide application in the trade. It is used in rubber as a toughening agent; in printers’ ink as a pigment; in paints; as a black coloring matter. Bone black is an amorphous carbon produced by destructive distillation of bones. It is used as a deodorizing and decolorizing agent.

THE MANUFACTURE OF ILLUMINATING GAS

Cylindrical ovens of fire clay (B—Figure 31) are filled automatically with soft coal. These ovens are then closed tightly to prevent entrance of air. Under these over, a hot fire (A) is maintained. The heat decomposes the coal into gases, liquids and coke. The gases contain impurities such as hydrogen sulphide, carbon dioxide, ammonia, tar, benzol and water. These impurities are removed before the gases are passed into the receiving tanks.

The first step towards purification of the gases is as follows: The gases pass through the oven in 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 then 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, coal tar and loose sawdust. A spray of water is introduced at the top of the scrubber and the material filling it is thus kept moist. The remaining water and ammonia salts are then removed as the gases pass on to the purifier.

Fourth, the gases are passed through the “purifier” (FF), a rectangular box in which layers of quick lime, which absorbs water, carbon dioxide and hydrogen sulphide, are placed. After the moisture is removed, the gases, which now consist of hydrogen, nitrogen, marsh gas, olefiant gas, acetylene and carbon monoxide, are delivered immediately into the large gas tanks (G). These tanks are constructed in telescopic fashion 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 consumer. (H) is the entrance pipe. (I) is the exit pipe. (K) is the flue or chimney for the fire.

You may have noticed at times the tremendous flame which shoots up in the vicinity of gas tanks. This is especially especially noticeable at night. The cause of this seems to be mysterious. It is exceedingly simple. When live steam is passed over white hot coke away from air, in an oven for instance, water gas is formed. Water gas cannot be used for lighting purposes alone; it is mixed with the coal gas. Water gas consists of a mixture of hydrogen and carbon monoxide and is produced by the action of steam on hot carbon.

After the steam has been passed into the ovens for a time new coke must be added; in other words, the ovens must be charged regularly. Before emptying the ovens, valves shut off the gas connection of the oven with the rest of the plant. In the ovens there still remains a good deal of gas. This gas must be removed before the ovens are emptied of the coke. The quickest and best way is to burn it away. This is done, and the gas disappears in a tremendous flame. Thus the “aurora borealis” of the city is a mystery no more.

EXPERIMENT 262—Preparation of charcoal—pyro-ligneous acid

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 an alcohol lamp or gas flame. Notice that the paper turns red, proving that an acid is evolved. This acid is called pyro-ligneous acid and is essentially acetic acid. Acetic acid is also present in vinegar. Here it is formed in the process of fermentation resulting from the action of bacteria on sugar.

Now insert the 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 that is, it passes off as a vapor. The tar oils are the heavy resinous oils, brown in color, such as you may see at the bottom of your test tube.

To summarize, when coal is heated without admission 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 sulfone 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 zepic action of this 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 a distinctive taste which is characteristic of smoked foods.
Meats can be cured over night, while in the smokehouse several days are required. This curing is a rather new method and it is practiced except experimentally. The ham is painted with pyro-lignaceous 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 as well.

**EXPERIMENT 263—Decolorizing properties of charcoal**

Make a solution of potassium permanganate by dissolving a crystal in a test tube half full of water. Shake until all is dissolved and the solution is colored purple.

Now put into this solution three measures of powdered charcoal and, closing the mouth of the tube with the thumb, shake vigorously for two or three minutes. Now filter this solution in the funnel and notice the color of the liquid that runs through. It is nearly colorless, showing that charcoal has the property of absorbing color from certain substances.

Repeat this experiment with different colored solutions and notice which of the colors are absorbed by charcoal.

**EXPERIMENT 264—Decolorizing properties of charcoal**

Prepare some hydrogen sulfide water as shown in a previous experiment and notice the odor. To this solution add three measures of powdered charcoal, close the mouth of the test tube with your thumb and shake the test tube for three or four minutes. Now filter this solution and smell the liquid which runs through. Notice that the odor has been removed by the charcoal.

**EXPERIMENT 265—Absorbing properties of charcoal**

Dissolve a small piece of quinine pill about the size of a pin head in a test tube half full of water and taste the solution. Notice that it is bitter.

Now add three measures of powdered charcoal, close the mouth of the test tube with your thumb and shake for three or four minutes. Filter this solution and smell a little of the liquid which runs through. Notice that the bitter taste is practically gone.

Charcoal then removes colors, odors and tastes from certain solutions.

**ABSORBENT CHARCOAL**

**EXPERIMENT 266—Absorbing coloring matter with charcoal**

Put about ½ measure of cochineal in a test tube ¼ full of water. Warm the test tube for a few moments until the cochineal dissolves, forming red solution.

Now put two measures of powdered charcoal into the cochineal solution and close the mouth of the test tube with your thumb and shake three or four minutes. Filter the solution after you have shaken it thoroughly in order to separate the charcoal and you will find that the color will be a great deal lighter than it was in the original solution. By repeating this process several times it will be possible to get practically all of the color out of solution.

Try this experiment on logwood solution, or solutions containing little magenta dyes, or on any other colored liquids.

**EXPERIMENT 267—Decolorizing vinegar**

Fill a test tube one-quarter full of vinegar which has a brownish or yellow color. 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 268—Absorbing bitter taste with charcoal**

If you can obtain a small amount of quinine you will find that charcoal will absorb the bitter taste. Fill a test tube one-quarter full of water and add a very small amount of quinine, about the size of an ordinary pinhead. Shake this with the water and taste a drop of the solution and notice the bitter taste.

**EXPERIMENT 269—Absorbing charcoal from ground coconut shells**

Grind up a piece of dry coconut shell and thoroughly bake the material in a copper oven. This should be thoroughly carbonized by this treatment. After baking then 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 270—Absorbing charcoal from butternut shells**

Grind some butternuts and eat 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 271—Preparation of absorbent charcoal from kicory nut shells**

Follow same directions as given for Experiment 270.

**EXPERIMENT 272—Preparation of absorbent charcoal from white birch wood**

Carbonize some small pieces of white birch wood 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 273—Preparation of absorbent charcoal from pine wood**

Follow same directions as given for Experiment 272.

**EXPERIMENT 274—Preparation of absorbent charcoal from maple wood**

Follow same directions as given for Experiment 272.

**EXPERIMENT 275—Preparation of absorbent charcoal from chestnut wood**

Follow same directions as given for Experiment 272.

**EXPERIMENT 276—Surface tension and the rubber band**

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 into 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.

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 277—Changing the specific gravity of charcoal**

Wood charcoal floats in 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. Boil the water for several minutes, then remove the weight from the charcoal, and you will find that it no longer floats. This is due to the fact that the air was driven away from the pores of the charcoal by boiling. This experiment illustrates why wood becomes waterlogged and does not float.
CARBON DIOXIDE OR CARBONIC ACID GAS

When carbon and any combustible compound of carbon is burned, the carbon is converted into carbon dioxide as the final product of oxidation. This is a gas 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.

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 the carbon dioxide in the lungs is responsible for stimulating the respiratory center in the process of breathing. The deep and rapid breathing from excessive 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 in 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 anesthetic 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.

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 of a life-saver and alleviator of pain. We 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 asphyxia, physical shock, and partial drowning. Machines for administering mixtures of oxygen and carbon dioxide have proved to be of great value in these types of cases than those which use oxygen alone. The breathing-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 longer needed, is a general hospital practice.

CARBONATES OF METALS

The carbonates of many metals are insoluble in water and can be precipitated from solutions by means of a soluble carbonate such as sodium carbonate. In the following experiments you will obtain many colors. Filter off your insoluble precipitates and note the color of the carbonate salts.

EXPERIMENT 278—Strontium carbonate

Dissolve one measure of strontium chloride in a test tube half full of water. Then add one-quarter of water. Dissolve one measure of sodium carbonate and a heavy white precipitate of strontium carbonate is obtained.

EXPERIMENT 279—Nickel carbonate

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 half full of water. Now add the sodium carbonate solution a little at a time to the nickel sulphate. A light green precipitate of nickel carbonate will be formed.

EXPERIMENT 280—Zinc carbonate

Put a small piece of zinc metal in a test tube, add one measure of sodium bisulphate and fill the tube half full of water. Warm this solution and let it stand for three or four minutes so that some of the zinc metal will dissolve

Now prepare a solution of one measure of sodium carbonate in a test tube half full of water and add some of this solution to the zinc solution. A white precipitate of zinc carbonate will be formed which gradually settles to the bottom of the tube.

EXPERIMENT 281—Aluminium carbonate

Dissolve one measure of sodium carbonate in a test tube half full of water. Dissolve one measure of sodium sulphate in another test tube half full of water. Add the sodium carbonate solution to the sodium sulphate solution a little at a time.

Notice that at first an effervescence takes place and then a gelatinous precipitate of basic aluminium carbonate is formed.

EXPERIMENT 282—Chromium carbonate

Dissolve one measure of chrome alum in a test tube half full of water. Add a few drops of sodium carbonate solution and heat. A blue precipitate of chromic carbonate forms.

EXPERIMENT 283—Ferrous carbonate

Dissolve one measure of ferrous ammonium sulphate in a test tube half full of water. To this solution add a few drops of a solution of one measure of sodium carbonate in a test tube half full of water. A greenish precipitate of ferrous carbonate will form.

EXPERIMENT 284—Cobalt carbonate

Repeat the experiment using one measure of cobalt chloride in place of the ferrous ammonium sulphate.

EXPERIMENT 285—Calcium carbonate

Put one measure of calcium chloride in a test tube half full of water. Shake thoroughly and then add a few drops of sodium carbonate solution prepared as in the preceding experiment. A white precipitate will be formed which is calcium carbonate.

EXPERIMENT 286—Copper carbonate

Place one measure of azurite, which can be purchased at your drug store, and one measure of sodium bisulphate in a test tube. Fill the tube half full of water and shake for a few moments until the solution becomes perfectly clear. Now add a few drops of sodium carbonate solution prepared as before, and you will obtain a blue precipitate of copper carbonate will be formed.

EXPERIMENT 287—Manganese carbonate

Dissolve one measure of manganese sulphate in a test tube half full of water. Add a few drops of sodium carbonate solution prepared as before and you will obtain a white precipitate of manganese carbonate.
EXPERIMENT 288—Magnesium carbonate

Add one measure of magnesium sulphate to a test tube half full of water and dissolve by shaking. Now add a few drops of sodium carbonate solution, as before, and a white precipitate of magnesium 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.

With the return of cold weather, the spectre of carbon monoxide poisoning haunts every automobile driver. Unless humanity has exercised an excess of caution, we may expect occasional news items about persons warming up their motors in closed garages, being overcome by this insidious gas. Carbon monoxide is a product of imperfect combustion. When a fuel is burned under perfect conditions, carbon monoxide is not produced. The products of perfect combustion are carbon dioxide and water. Imperfect combustion conditions are most dangerous. Certainly not in the very best of automobiles. An analysis of the exhaust gases of an average automobile shows about 70% of carbon dioxide, 10% of carbon monoxide. In certain experiments a dog was left in the driver’s seat of an automobile in a closed garage, with the engine running slowly. Twenty minutes later the dog was unconscious. Had a man been in the dog’s place, the result would probably have been the same. Carbon monoxide overcomes its victim with no warning. The first symptom is a severe pain in the back of the head, but if the concentration is high, the victim may lose consciousness before he can act on this warning. This condition may or may not be preceded by such warning symptoms as headache, dizziness, or nausea. Small doses may have no other effect than to cause severe headache, but heavy gassing is a serious matter. Convulsions from carbon monoxide poisoning should be kept in bed even when they protest that they are all right. To avoid carbon monoxide poisoning it is not enough to good ventilation in the garage. Carbon monoxide is lighter than air and vanishes immediately through an open window or door. A doctor should be called immediately for a person overcome by gas. An automobile driven by a driver under the influence of carbon monoxide gas constitutes a hazard to the public safety on the highway equal in seriousness to that of another car equipped with faulty brakes. It has been found that an automobile, following closely behind another, particularly in heavy traffic, may pick up a sufficient quantity of the exhaust gas from the preceding car to result in a dangerous mixture with the second automobile, leading to carbon monoxide poisoning.

ETHYL GAS

Every automobile driver is familiar today with the trade term—"ethyl gas." This is an organic compound containing lead which hastens on gasoline the favorable properties characteristic of this reagent (tetrachlorothane). It is a practical anti-knock substance. When ethyl gasoline is used, it tones down the explosions in the cylinder of gas engines, and makes the cylinder more gently than in the case when ordinary gasoline is used. Tetrachlorothane 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. Ordinary precautions are followed, tetrachlorothane gasoline presents little danger during

EXPERIMENT 289—Preparation of carbon dioxide—(effervescence)

Make a solution of sodium bicarbonate by shaking up a test tube 1/4 full of water containing two measures of sodium bicarbonate. In another test tube make a solution of tartaric acid by shaking up a test tube 1/4 full of water containing two measures of tartaric acid. Now pour the tartaric acid solution into the solution of sodium bicarbonate. 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 290—Vinegar and baking soda

Obtain some very strong cider vinegar, the stronger the better, and place about 3 cc in a test tube. Then drop into the vinegar a small measure of baking soda. What gas is given off? Test for it.

EXPERIMENT 291—Vinegar and oyster shells

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.

EXPERIMENT 291—Vinegar and oyster shells

Repeat the above experiment using some painter’s thinning. What is the composition of ordinary thinning?

EXPERIMENT 291—Vinegar and chalk

Repeat the above experiment using some powdered chalk or crayon from your schoolroom blackboard.

EXPERIMENT 291—Vinegar and marble

Repeat the above experiment using some pulverized marble.

EXPERIMENT 295—Vinegar and tooth powder

Repeat the above experiment with a good quality of tooth powder. What constituent of tooth powder causes the reaction?

EXPERIMENT 296—Vinegar and old mortar

Repeat the above experiment using some old mortar removed from the walls of an old brick building. Note the evolution of carbon dioxide gas.

EXPERIMENT 297—Vinegar and Portland cement

Repeat the above experiment using some pure Portland cement.

EXPERIMENT 298—Carbon dioxide is heavier than air, and will not burn

Light a candle and set it firmly on a board by sticking it to a little melted wax from the flame of the candle. (Figure 32).

Now put one-half teaspoonful of sodium bicarbonate or 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 over the flame just as though you were pouring water out of the glass, taking care that carbon dioxide is heavier than air and will settle to the earth and also that it will not burn.
EXPERIMENT 299—Chemistry of the flame

Examine closely an alcohol lamp flame or candle flame and observe that it consists essentially of three cones. (Figure 33.) First a dark cone just around the wick; second, a yellow cone which produces light; and third, a semitransparent cone of heat around the outside.

The dark cone of the flame consists of unburned gases which are given off from the wick of the candle. The paraffin is melted by the heat and drawn into the wick by capillary action. In the wick the paraffin is converted into a gas by the heat of the flame. To prove this, hold one end of a hollow glass tube in the flame just over the wick. Now apply a flame to the other end of the tube and it takes fire. The gas in this cone is relatively cool, for if a match stick is placed in it that portion of the stick which was in the dark cone will not burn as soon as the portion passing through the sides of the flame.

The second or yellow cone of the flame consists of particles of carbon that have been heated to white heat so that they glow brightly.

Hold a cold spoon or glass rod in this cone for a minute and notice that it is covered with a black deposit of carbon called lampblack, thus proving that this cone consists of small particles of carbon. The cold spoon chills the flame, thereby causing the carbon particles to be deposited. Lampblack is made upon this principle on a manufacturing scale.

The third or outer cone of the flame consists of the gases formed by the complete burning of the carbon particles to carbon dioxide gas. This is the hottest portion of the flame, and whenever heating a liquid in a test tube, for example, it is important, in order to obtain the highest heat possible and to prevent the deposit of soot, to hold the test tube at the tip of the luminous or light-giving part of the flame.

EXPERIMENT 300—Carbon dioxide from a burning candle

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.

EXPERIMENT 301—The structure of a flame—A gas factory

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.