the danger of using hydrogen as a buoyant gas in airships. A leak in a gas bag is sufficient to cause the destruction of a dirigible. In order to avoid this danger, air or helium is now substituted for hydrogen, and is undoubtedly the only safe gas for use in airships, or in inflating balloons. The discovery of helium is really one of the most significant events in the history of science. The United States is fortunate in having supplies of natural gas and helium in the State of Texas and in other parts of this country which contain sufficient helium and natural gas to allow of its separation on a commercial scale. Helium, discovered previously in association with the rare gases of the atmosphere, is surpassed only by hydrogen in lightness of weight, but it is superior to hydrogen in that 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 buoyant power.

<table>
<thead>
<tr>
<th>Gas</th>
<th>Specific gravity at 0°C and 760 mm Hg</th>
</tr>
</thead>
<tbody>
<tr>
<td>Air</td>
<td>1.000</td>
</tr>
<tr>
<td>Hydrogen</td>
<td>0.0894</td>
</tr>
<tr>
<td>Helium</td>
<td>0.1380</td>
</tr>
<tr>
<td>Nitrogen</td>
<td>0.9672</td>
</tr>
<tr>
<td>Oxygen</td>
<td>1.1037</td>
</tr>
</tbody>
</table>

**EXPERIMENT 68—Preparation of hydrogen**

Mix together in a test tube 3 measures of sodium bisulphate and 4 measures of ammonium chloride and add about one inch of water. Warm gently until the mixture completely dissolves.

Then add 1 measure of powdered zinc and notice the violent reaction that takes place. This is due to the evolution of hydrogen gas. Hold a flame at the mouth of the test tube and notice that the gas given off burns with a blue flame. This flame is footed by hydrogen combining with the oxygen of the air to form water.

Hydrogen and oxygen form a very explosive mixture, and great caution must be exercised when hydrogen gas is brought in contact with oxygen in large volumes.

In the reaction above the sodium bisulphate and ammonium chloride interact and form hydrochloric acid, which then reacts with the zinc powder to form a soluble zinc chloride and hydrogen gas.

**REDUCTION**

Hydrogen can be heated at a high temperature in the absence of oxygen with impunity. It interacts with many substances under specific conditions with 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 on oxygen 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 and if we need to remove oxygen to accomplish oxidation. Hydrogen peroxide and hydrochloric acid are chemical reagents which react as oxidizing agents.

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

**BLEACHING**

Bleaching is a commercial term widely used in the 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 tussah silk. The gas sulphur dioxide is a cheap, commercial reagent which is a valuable bleaching agent and extensively used in industry, especially in the linoleum and straw trade.

**EXPERIMENT 71—Bleaching flowers with sulphur dioxide**

Place about 6 measures of sodium bisulphite and an equal amount of tartaric acid in a glass tumbler (figure 22), and add a quarter test tube of water. An intense reaction sets in, sulphur dioxide is liberated. Notice the color of the gas.

**EXPERIMENT 69—Reduction of potassium permanganate**

Place 2 small crystals of potassium permanganate in a test tube half full of water and shake well to dissolve the crystals. Notice the purple color of the solution. Now add 3 measures of sodium bisulphite and shake again. Notice that the color of the permanganate solution is destroyed. Here we have to deal with a reaction between an oxidizing and a reducing agent. The two are incompatible. The sodium bisulphite has power to take up oxygen donated by the potassium permanganate, thus converting it into a colorless substance. We may say that the potassium permanganate has oxidized the sodium bisulphite, or that the sodium bisulphite has reduced potassium permanganate. In other words, the experiment serves to illustrate both an oxidation and a reduction process.
Moisten some red or blue flowers with cold water, put them in the tumbler, the glass with a saucer, and allow to stand. After a period of one hour the flowers be bleached to a pure white. The generated sulphur dioxide has reduced the natural coloring substances of the flowers and destroyed them.

**NITROGEN**

Nitrogen is a colorless, tasteless gas. It forms four-fifths of the bulk of the air. Chemically speaking, this is a very inactive gas, 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 62 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 the 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 the compounds would be unstable and would decompose easily. This is true, and of the high explosives used during military operations today are nitrogen compounds.

Many of the most highly explosive substances are compounds of nitrogen, oxygen, and sulfur. The force of an explosion is due to the tremendous volume of gases that are simultaneously formed when an explosive compound is detonated. Cannon or a shell. Gunpowder, nitroglycerine, nitrokabole, or gun cotton, picric, trinitrotoluene (T.N.T.) and tetranitramine (T.N.A.) are some very common 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 substance of protein, this element stands among the first of the 93 elements in every-day importance. Because nitrogen is needed by the animal and plant life, it is a constant source of supply of this element for growing plants, animals, and man. Plants obtain their nitrogen from the soil, soil nitrogen being built up by the action of bacteria which are capable of converting nitrogen compounds into nitrate or by the action of fertilizers containing nitrogen. Nitrogen also features in other industries, among them lacquer, coated textiles and plastics.

Nitrogen combines with oxygen to form several oxides. Nitric acid (HNO₃) compound of nitrogen with oxygen is the most important acid in 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 large scale by treating chlorate with sulphuric acid. This acid, along with ammonia, are manufactured today in enormous quantities from nitrogen and oxygen of the atmosphere.

**EXPERIMENT 72—Preparation of nitric acid**

Put 4 measures of potassium nitrate and 4 measures of sodium bisulphate in a 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 the alcohol lamp. Notice that the blue litmus paper turns red. Remove the test tube from the flame and notice the smell of ammonia in the air with the test tube and notice the change in the paper. It will turn blue. Any ammonia salt when heated in the presence of acid or alkaline (lime water, sodium carbonate, and sodium hydroxide) will give off ammonia gas and this reaction is used as a test for the ammonium group.

**AMMONIA**

Any 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 ammonia is used for cleaning purposes and for softening. Ammonium carbonate, a salt which readily gives up its ammonia fumes under ordinary temperatures, is commonly used as smelling or aromatic salts. Ammonia gas is soluble in water, and the aqueous solution—ammonia water—is the form in which it is commonly met with in the trade. Ammonia gas is condensed and shipped in enormous quantities today in iron cylinders. Ammonia in this form finds a wide use in artificial refrigeration. Formerly most of the ammonia of commerce was obtained from the destructive distillation of coal. It is a valuable by-product in the manufacture of nitric acid gas. Today the gas is made in large quantities synthetically by combining nitrogen of the air with hydrogen.

**EXPERIMENT 75—Preparation of ammonia**

Put 2 measures of sodium carbonate and 2 measures of ammonium chloride in a tube. Add 1 spoonful of water and heat gently. Remove the test tube from the flame and notice the smell of ammonia in the air with the test tube and notice the change in the paper. It will turn blue. Any ammonia salt when heated in the presence of acid or alkaline (lime water, sodium carbonate, and sodium hydroxide) will give off ammonia gas and this reaction is used as a test for the ammonium group.
(NH₄) in a compound. If the gas produced by heating the above mixture is introduced into cold water, it is dissolved and ammonium hydroxide is formed.

**EXPERIMENT 76—Formation of ammonia by decomposition of organic matter**

Most organic matter contains nitrogen and when heated in the presence of an organic base, like calcium oxide, the nitrogen is partially liberated in the form of ammonia (NH₃). Ammonia was first prepared by this method and was called "Spiritus Hartzhorn."

Place a small quantity of wool, hair, silk or finger nail clippings in a test tube, add 3 measures of calcium oxide and 3 drops of water. Place a small strip of cored red litmus paper over the mouth of the tube and gently heated the tube over a flame. Notice that the red litmus paper turns blue, showing that a volatile compound has been formed. Remove the tube from the flame and smell at the mouth of the tube. We recognize the odor of ammonia if the experiment is properly conducted.

**EXPERIMENT 77—Dissociation of an ammonium salt**

Put one measure of ammonium chloride in a clean dry test tube and heat over a flame. Notice that the dry salt (NH₄Cl) passes into the vapor state and condenses again on the cooler part of the tube to form again solid ammonium chloride. What really takes place when heat is applied is that the ammonium salt is split up into molecules of ammonia (NH₃) and hydrochloric acid (HCl), which combine in the cool part of the tube to form again solid ammonium chloride.

**WATER**

At ordinary temperatures pure water is a tasteless, odorless, transparent liquid unless in thin layers, but distinctly blue when seen in large masses. It is always heavier than air.

Water consists of two elements—both gases—hydrogen and oxygen, and thus in the proportion of 89 per cent of oxygen and 11 per cent of hydrogen by weight and volume, water consists of two parts of hydrogen to one part of oxygen.

Water occurs very abundantly throughout the earth. Vast areas of the cold pole of the earth are covered with it in the form of ice, while in the liquid state it is present in about five-sevenths of the earth's surface, residing in some places to a depth of six miles. Large quantities occur in the soil, and as a vapor it is an essential part of the atmosphere. More than half the weight of living organisms consists of water. Great many substances dissolve in water, so that water is known as a very universal solvent. A substance is said to be dissolved when none of the particles of the substance can usually be recovered from the 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 in the roots is given off by the leaves. The stream thus maintained serves the plant cells fully distended thereby preserving the form of the plant, and it is drinking to carry on its vital processes. Even though the largest proportion of it is used in the growing life of a plant is given off through its leaves, a constant amount is retained. In succulent plants, about 90 per cent of the complete water content of the plant is water, this water being an actual part of the organism. In wood, the percentage of water present in the organism is less, although even in this find an important compound an essential part of the living cell. With animal tissue, it is as essential as it is with plants. About three-fourths of the total weight of the human organism is water, and about two quarts of drinking water a day, best.
DISSOCIATION OF WATER INTO THE ELEMENTS OXYGEN AND HYDROGEN

EXPERIMENT 78—Decomposition of water by an electric current (electrolysis)

It can be definitely proven that water is composed of two gases—namely, oxygen and hydrogen. 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 arc light carbon or clean lead.

The apparatus is set up as shown in the illustration (Figure 23). 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 ends of the copper wires and the two cell ends, attach the long pieces of copper wires and fasten to the free ends of the two gas pieces of arc light carbon or lead, having previously scraped the free ends of the carbon wire with a knife to insure a clean surface.

Now fill the pan about half full of a solution of sodium bisulphate, using one teaspoonful of the compound to each glass of water required. Fill two test tubes with sodium bisulphate and invert them upside down in the solution, being careful not to allow any bubbles of air to get into them.

\[ \text{Fig. 23} \]

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EXPLANATION OF THE ELECTROLYSIS EXPERIMENT

Now put one of the carbon or lead electrodes under each test tube, being careful not to allow any air in the tubes. Notice that immediately bubbles of gas begin to form. The gas formed in each test tube from the carbon or lead electrodes and gradually force their way to the surface of the water. 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 electrode or cathode, and the electrode at which the oxygen is liberated is the positive electrode or anode.

AQUEOUS SOLUTION

EXPERIMENT 79—A soluble substance

Dissolve one or two teaspoonfuls of common salt in a glass of water. Notice that the solution is not clear. If this liquid was filtered, you would find that nothing was left in the filter paper.

EXPERIMENT 80—An insoluble substance

Dissolve one or two measures of sulphur in a test-tube half full of water. Well and notice that the sulphur remains in the solid state. Place the solution into a clean pan and heat over a stove until all the water is driven off. The sulphur remains. Allow this to cool and notice a little of the same salt you dissolved in the water, proving that the salt when in solution undergoes a physical change only.

EXPERIMENT 81—Diffusion

Drop a grain of potassium permanganate in a glass full of water and notice what happens. The solid dissolves, you will see the red color traveling up through the water and after a few minutes the water will be of a uniform color, due to the complete diffusion of the potassium permanganate throughout the liquid.

EXPERIMENT 82—Diffusion of cobalt in water glass solution

Dissolve one tablespoonful of syrupy water glass solution in a tall tumbler full of water. Place the tumbler on a table where it will not be disturbed and drop into the solution three or four crystals of cobalt chloride. A column of cobalt chloride will grow down out of the mouth of the tube. Also notice that there is twice as much cobalt chloride growing from the bottom.
EXPERIMENT 83—A chemical garden

Take a two-quart fruit jar with a wide mouth and fill with a previously solution of water glass. Then plant in the bottom of the jar crystals of the salts in the following order: copper sulphate and nickel ammonium sulphate, follow after about thirty hours with a planting of crystals of magnesium sulphate and manganese sulphate. Finally after the diffusion growth is well along, add the salts of cobalt chloride. A beautiful chemical garden effect will be produced.

EXPERIMENT 84—Diffusion through cellophane

Wrap some crystals of copper sulphate in a tight cellophane bag. Be careful crack the cellophane bag by folding. Suspend the bag containing copper sulphate in a jar of water, being careful that no water gets inside the bag. Water will diffuse through the cellophane film and dissolve some of the copper sulphate. The copper sulphate will slowly pass through the cellophane and color the outside solution. Passage of a salt through a membrane like cellophane or parchment is called diffusion.

EXPERIMENT 85—Separation of starch from sugar

Prepare a mixed solution of sugar and starch and transfer to a cellophane bag. Suspend in a jar of water and sugar will diffuse through the cellophane bag into the outside liquid. Taste the water after a few hours and you can the sugar by the taste. Starch will not diffuse through the cellophane bag into the solution. To prove this, test some of the outside solution with starchiodine paper. Starch is present; a blue coloration will be produced. By changing the outside container, several times all the sugar can be separated from the starch in the cellophane membrane.

EXPERIMENT 86—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.

CHEMICAL ACTION IN WATER SOLUTION

Most chemicals, as a rule, do not interact when brought together in the dry state. 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 by solution in water the molecules are changed or activated, and the particles of compounds are brought much closer together. Other solvents besides water are used as a solvent to promote chemical change. For example, liquid ammonia serve as a solvent for carrying out many interesting reactions that cannot be done in water solutions. Alcohols, acetone, benzeno and other organic reagents are placed of water as solvents for carrying out organic reactions. The organic chemist inorganic chemist apply an entirely different technique in accomplishing reactions.

EXPERIMENT 87—Promoting chemical reaction by solution

Mix together on a sheet of paper one-half spoonful of sodium bicarbonate and half spoonful of tartaric acid. Notice that there is no reaction. Transfer this mixture to a test tube and add a few drops of water. A violent results with the liberation of carbon dioxide gas, thereby proving that water necessary to promote chemical reaction between these substances.

EXPERIMENT 88—A color change due to solution

Take a mixture of one measure of tannic acid and one measure of ferric ammonium sulfate on a piece of paper and notice that there is no evidence of reaction. Transfer to a clean, dry test tube and fill half full of water. Notice the formation of black colored product, showing that there was a chemical reaction due to the influence of water. The black product formed is iron tannate. This black substance is a constituent of many inks.

DRI ICE EXPERIMENTS

EXPERIMENT 89—Extending a burning candle

Place a burning candle in the bottom of a pint fruit jar. Then drop a spoonful of dry ice on the bottom of the jar. Notice what happens when the dry ice turns to gas. What do you explain this?

EXPERIMENT 90—Burning sulphur

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

EXPERIMENT 91—Burning magnesium

Repeat the above experiment by letting the dry ice vaporize and drive all the air the fruit jar. Then drop a small quantity of powdered magnesium into the jar of dry ice and liberating carbon dioxide gas. The magnesium will continue to burn, using the oxygen of the dry ice and liberating particles of finely divided carbon. Magnesium has greater for oxygen than carbon.

EXPERIMENT 92—Floating a sunken ship

Put a spoonful of dry ice in a thin rubber bag and tie with a silk thread so that it will not escape. Then attach a weight to the bag so that it will sink to the bottom of the jar of water. Note the result as the solid dry ice gasifies inside the rubber bag. The bag will finally float. How do you explain this?

EXPERIMENT 93—Measuring dry ice pressure

Take a U-tube made of glass tubing about 5 mm. diameter and fill the lower curve with a little mercury. Then drop some dry ice into one arm of the tube with close the mercury in the U-tube immediately with a tight rubber stopper. Observe what happens when the second arm of the tube and finally be expelled if the arm is of small diameter.

EXPERIMENT 94—A miniature gas volcano

Place a heavy wall tube and fill with dry ice, then place in upright position in the bottom of a tall glass jar filled with water. As the dry ice vaporizes the gas through the openings of the salt seller lid and the bubbles will travel to the outlet of the water. Hold an electric bulb back of the jar of water during the bubbling.

EXPERIMENT 95—Dry ice temperature

Place a piece of dry ice in an ordinary glass tumbler. Read the temperature with your thermometer. Now pour into the tumbler one test tube of strong lye solution. Stir the ice. Notice the change that takes place and observe the change in temperature. The dry ice is a weak acid and will slowly dissolve in alkali. After the dry ice has dissolved then add some vinegar to the alkaline solution until limewater paper turns red.
ENDOTHERMIC AND EXOTHERMIC CHANGES

Some compounds give off heat when they dissolve in water. This is an exothermic change. Other compounds absorb or take up heat when they dissolve. This is an endothermic change.

EXPERIMENT 96—Lowering of temperature by solution

Fill a test tube half full of water, and notice the temperature by feeling its tube with the face or hand. Read the temperature by means of your thermometer, giving each about the same amount of heat. Notice the change of temperature by means of your thermometer. Compounds, like ammonium chloride, sodium nitrate and many others have, as is known, a negative heat of solution. (endothermic) That is, they absorb heat of water when dissolved in it, thereby lowering the temperature or cooling the water.

EXPERIMENT 97—Raising of temperature by solution

Fill a test tube half full of water and notice the temperature by feeling with the face or hand. Now dissolve in the water half a spoonful of magnesium sulphate. Notice the temperature the same as before. This time the solution is warmer. Compounds like magnesium sulphate, sodium hydroxide, and some others have a positive heat of solution (exothermic). That is, they give off heat when dissolved in water.

ARTIFICIAL REFRIGERATION

The application of the so-called negative heat of solution of common salt to the use of it as an 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 salt and water whose temperature is just at the freezing point, the salt dissolved in it lowers the temperature of the solution several degrees below the freezing point of water thereby affording us a very convenient freezing mixture.

EXPERIMENT 98—How to make a freezing mixture

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 it to stand for several minutes. Notice that the water in the test tube will freeze to solid ice after a short time. Use your thermometer for noting the changes in temperature.

CHANGING THE FREEZING AND BOILING POINT

It is easily demonstrated that when a substance is dissolved in water the freezing point of the resulting solution is lower than the freezing point of water. The reason why salt is sometimes thrown on slippery sidewalks. It melts the ice and lowers 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 99—Solution lowers the freezing point of water

Make a freezing mixture as shown in the preceding experiment. Now add a full of common salt to a test tube half full of water and shake until dissolved. Take a test tube of salt solution with another test tube half full of water in the freezing place and allow to stand. Note the change in temperature by means of a thermometer. Notice that the water freezes but that the salt solution does not, thereby proving that solution lowers the freezing point of water.

EXPERIMENT 100—Solution raises the boiling point of water

Take 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 an alcohol lamp and allow to boil. Notice the temperature the same as before. This time the solution is warmer. The salt solution raises the boiling point of water. Note the change in temperature by means of a thermometer.

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 remain a liquid. When in this state, water is said to be undercooled.

EXPERIMENT 101—UNDERCOOLED WATER

Take your freezing mixture as explained in previous experiments. Place in the freezing mixture one test tube one-third full of water and keep the test tube quiet. The temperature of the water in the tube may go down as far as 8 to 10 degrees below the freezing point 0 degrees Centigrade or between 18 and 14 degrees Fahrenheit, and the water 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.

DEGREE OF SOLUBILITY IN WATER

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

EXPERIMENT 102—Effect of temperature on solubility

Take seven measures of nickel ammonium sulphate in a test tube one-quarter full of cold water and shake well. Notice that some of the solid remains undissolved. Now heat the test tube slowly and notice that all the solid goes into solution, showing that substances are more soluble in hot water than in cold water.

Try the effect of this mixture upon another test tube half 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. Notice the liquid gets cloudy or turbid, proving that calcium hydroxide which was formed when calcium oxide was added to the water is less soluble in hot water than in cold water.

EXPERIMENT 103—Temperatures Effect on Solubility

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

EXPERIMENT 104—Dissolving liquids in each other

Try to mix a few drops of carbon tetrachloride with a little water in a test tube. You can shake as much as you wish but the carbon tetrachloride will always settle on the bottom of the test tube. In other words, it is insoluble in water.
Try dissolving a few drops of glycerine in a little water in a test tube. In this way the glycerine soon mixes and stays in solution with the water.

**EXPERIMENT 105—Dissolving a gas in water**

Put two measures of sodium carbonate and two measures of sodium bisulphate in a test tube and fill the tube one quarter full of water. Quickly fit the test tube with a gas delivery tube and stopper and insert the other end of the delivery tube into the mouth of an empty test tube so that the gas which is coming from the delivery tube will flow into the test tube just as if water were coming through the delivery tube instead of gas which you cannot see.

After a minute or so the test tube should be full of gas. Keep your finger over the mouth as much as possible and carefully fill the tube one half full of water. Press your finger or thumb tightly over the mouth and shake. Do you feel a pull on your thumb showing that the gas in the tube has dissolved in the water by in a vacuum?

**EXPERIMENT 106—Removing a gas from solution by boiling**

Put one drop of ammonium hydroxide in a test tube one quarter full of water. Now add one drop of phenolphthalein solution. This will color the liquid in the tube red, as ammonium hydroxide is an alkali. Heat this red liquid for a few minutes and notice that as it begins to boil the red color becomes fainter. By boiling for several minutes the liquid may be turned colorless, but this is rather difficult to do as there is a great tendency for the liquid to slop out of the tube when boiling.

Ammonium hydroxide is a gas dissolved in water and when heated this gas is driven out of the solution, leaving only water.

**EXPERIMENT 107—Diffusion**

Fill a clean glass nearly full of clear water and let it stand for a minute or two and it becomes quiet. Now add a small quantity of mixed dyes and watch closely what occurs. As the substance dissolves the color seems to flow out from the small crystals, gradually spreading over the bottom of the glass. After a few minutes the color will diffuse upward through the liquid until finally after several hours the entire solution will be a uniform color.

**EXPERIMENT 108—Testing for acidity**

Moisten a piece of blue litmus paper with a drop of the water you are testing. If the blue litmus paper turns red or pink, the water is slightly acid.

**EXPERIMENT 109—Testing for lime**

To a test tube of water add two measures of sodium carbonate and shake well. If the water, after standing a while, shows a white turbidity it contains a considerable portion of lime.

**EXPERIMENT 110—Testing for sulphates**

To a test tube full of water add one measure of strontium chloride and shake well. If the water after 10 minutes shows a white turbidity it contains a considerable portion of sulphates.

**EXPERIMENT 111—Testing for iron**

To a test tube full of water add one measure of sodium ferrocyanide and shake well. If the water shows a blue tinge either at once, or after standing a while, iron is present.
What you really did was to distill the copper sulphate solution. That is, the steam 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 non-volatile remains 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. Ferric sulphate, nickel sulphate, and copper sulphate, for example, contain water of crystallization.

Some substances give up or lose their water of crystallization by simple evaporation, and do not require the assistance of heat or air. Such substances are called "Deliquest's Substances" and a good example of this class of substance 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 "Deliquest's Substances," and a good example of this class of substance is calcium chloride.

**EXPERIMENT 116—Water of crystallization**

Pour two measures of copper sulphate in a clean, dry test tube and heat over a flame, using the test tube holder so as not to burn the fingers.

You will notice that water in the form of steam is given off and some of the water deposition on the inside of the test tube in the form of small drops of water. That which was originally in the copper sulphate is called water of crystallization.

Notice also that the salt remaining in the test tube has changed color from white. Many salts, when they lose their water of crystallization in this manner, also undergo a change in color. This dry, colorless compound is known as anhydrous copper sulphate, meaning without water of crystallization. Add four drops of water to the white copper sulphate and notice the change of color to blue.

Everybody at some time or other has eaten rock candy or crystallized sugar. This is nothing more than sugar that has been allowed to crystallize out from a concentrated solution of the sugar in water. A substance crystallizing slowly from solution tends to come out in the form of large, shapely crystals, while when crystallizing quickly from a solution it comes out in smaller crystals.

**EXPERIMENT 117—Formation of crystals (rock candy)**

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 of it and allow the contents of the tube to cool slowly undisturbed (Figure 23).

After a time the sugar will appear in large crystals upon the string which is hung from the test tube.

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

**EXPERIMENT 118—Disappearing ink**

In a small test tube make a solution of cobalt chloride by dissolving three measures of cobalt chloride in one-half inch of water. Shake thoroughly to dissolve all the solid. Procure a clean pen, or better a toothpick, write upon a pink paper with the pale solution you have made. Allow this to dry. When dry, the writing will be invisible. Go back to view the words you have written, merely hold the paper, taking care not to touch or burn it. The writing appears in a beautiful cobalt blue color.

Make the writing disappear again to allow to cool or hold it over a steaming kettle. Normally, cobalt chloride is blue. It has a great attractive power, however, for moisture. Therefore, in the presence of moisture it takes up what is known as "water of crystallization." The blue crystals then turn pink in color. Should the pink crystals become dry they would turn blue again, due to the loss of water.

When the cobalt chloride solution dried upon your paper, minute crystals were detached which contained water of crystallization. They were pink or colorless. When these tiny crystals turned blue because the water was driven off. When the salt dried, they again gathered moisture from the air and became colorless. Often in the weather is very clear or when the air is hot and dry, especially in winter, a long time is required for the blue color to fade.

**EXPERIMENT 119—How to make a weather barometer**

Ask a piece of unglazed paper in a concentrated solution of cobalt chloride and allow to dry. The solution of cobalt chloride is made up the same as in the preceding experiment, using three times the amounts.

When hung in the open air, the color of the paper indicates the state of the weather. It is well to notice that this is an indication of a dry, clear atmosphere or cold weather.

When many chemicals separate from their saturated solutions in hot water, they will cool slowly, in the form of crystals. Each substance that crystallizes has its peculiar crystalline formation. Large crystals are best obtained by allowing the solution to cool very slowly.

**EXPERIMENT 120—Crystals of copper sulphate**

Prepare a saturated solution of copper sulphate in water and pour the clear liquid into a small beaker. Then set aside to cool slowly, after placing a piece of cardboard over the top of the beaker to avoid rapid cooling. After the crystals have completely separated, then remove several of them and examine under your microscope.

**EXPERIMENT 121—Crystals of tartaric acid**

Repeat the preceding experiment using tartaric acid as your chemical.

**EXPERIMENT 122—Crystals of iron alum**

Repeat the above experiment using ferric ammonium sulphate as your chemical.

**EXPERIMENT 123—Crystals of magnesium sulphate**

Repeat the above experiment using magnesium sulphate or Epsom salts.
EXPERIMENT 124—Crystals of Glauber's salt

Repeat the above experiment using sodium sulphate. Preserve all the salt crystals in the air and then preserve the specimens for future experiments.

TESTING WATER

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

The principal impurities found in water are salt, and compounds of calcium, magnesium and iron. Waters containing such substances in solution are commonly known as hard waters, and, if large amounts of mineral matter are present as mineral waters. Some of the natural mineral waters possess valuable medicinal properties and consequently are set aside and protected for public use. Springs of mineral water are found in several parts of the United States. The salts occurring in these 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; quite often this is accompanied by certain forms of micro-organisms or living beings which may be injurious to life. Typhoid fever is quite often 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. Therefore, it is very important to make bacteriological tests of drinking water, especially in large cities. The government has passed laws in various states to prevent the sale of milk when the gold or the plate test is not passed. A similar law is passed in some cities. Sewage is also a menace to public health. The unthinking citizen believes that the problem of sewage disposal is one that is coming to the front because sewage improperly treated is a menace to public health. The unthinking citizen believes that the problem of sewage disposal is one that is coming to the front because sewage improperly treated is a menace to public health. The unthinking citizen believes that the problem of sewage disposal is one that is coming to the front because sewage improperly treated is a menace to public health.

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. Great prominence has been given them as causative agents of disease. Therefore, it is perfectly natural that bacteria in general would be looked upon as organisms harmful to man. This is not true. Taking the entire group as a whole, they are beneficial in that the good they do far outweighs the harm. While we cannot name the most important processes which depend on microbial activity, the greatest 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 important elements—which we could not do without—would remain locked up in 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 objects, there are good ones and bad ones, and great is the loss of life and property and suffering for which some bacteria are responsible, the beneficial effects outweigh the harmful effects.

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 precipitation in some form. There are two kinds of hard water—temporary and permanent.

EXPERIMENT 125—Temporary hardness—How to get rid of it

You are able to obtain some hard water in your locality, test a half test tube full of the water and heat it over a flame. If the water boils becomes turbid, that is, takes on a white milky color, it possesses temporary hardness.

Temporary hardness is due to the presence of calcium bicarbonate, which is formed by the action of lime (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 milky or milky appearance.

PERMANENT hardness is due to the presence of calcium carbonate and magnesium carbonate, which precipitate upon boiling. Kettles and boilers become gradually covered inside with a brown deposit, and in large scale it is a serious problem in manufacturing plants. To remove this deposit and the process, it is a good method to boil the water in such a manner as to allow the deposit to form slowly and then remove it by filtering.

EXPERIMENT 126—Permanent hardness—How to get rid of it

A sample of hard water in your locality, treat a half test tube of the water with 2 measures of sodium carbonate. Shake well and if a white precipitate is formed, the water possesses permanent hardness. This precipitate may be calcium carbonate or magnesium carbonate or both.

PERMANENT hardness is due to the presence of sulphate of lime or magnesium. The latter is known as permanent hardness. You are able to get rid of it by treating the water with enough lime to form a white precipitate. This water is safe for drinking purposes.

EXPERIMENT 127—Clarifying water

One-half measure of aluminium sulphate to a test tube three-quarters full of water prepared by adding a little clay to some muddy water and shaking. Then add one-half measure of sodium carbonate and shake again. Allow the tube to stand...