4

Solubility

4.1 EXPERIMENT DISSOLVING A SOLID IN WATER

You know from daily experience that sand and chalk do not dissolve in water, but that sugar and table salt do. Of course, these are only qualitative observations—that is, observations that do not involve measurement. Are you sure that not even a tiny amount of chalk dissolves in a gallon of water? Can you dissolve as much salt in a glass of water as you wish?

Many solutions are colorless—there is nothing to be seen once the solid is dissolved. To make things visible, we shall begin our quantitative study of solutions with a solution that has color.

Place 5.0 cm^3 of water in one test tube and 20.0 cm^3 of water in another. Add 0.30 g of orange solid to each tube, then stopper and shake thoroughly. **CAUTION:** Be careful not to get any of the solution on your hands. If some spills, wash thoroughly with water.

- Did all of the solid dissolve in both test tubes?
- Do you think that each cubic centimeter of solution contains the same mass of dissolved material?
- Both tubes contain the same mass of dissolved material, but is the shade of the color the same in both tubes?
- Will another 0.30 g of the orange solid dissolve in each of the test tubes? Try it, but be patient!

The solution may cool as the solid dissolves. If so, try to keep the temperature fairly constant by warming the test tube with your hand.

- Did 0.60 g of solid dissolve as well in 20.0 cm^3 as in 5.0 cm^3 of water?
- Is the color uniform in each solution?
- Is the shade of the color the same in both solutions?

Add another 0.30 g of the orange solid to each solution.

- How much do you have in each test tube?
- What do you observe?
- How much orange solid do you think there would have to be in the 20.0-cm^3 test tube so that not all of it will dissolve? Test your prediction.

A solution in which no more solid can be dissolved is called a "saturated" solution.

4.2 CONCENTRATION

The uniformity of color in any one of the various solutions in the last experiment suggests that the solid dissolved uniformly. That is, a cubic centimeter of water in a given solution contained the same mass of orange material as any other cubic centimeter of water in that solution. For example, at the beginning of the experiment one test tube had $0.30\ g/5.0\ cm^3 = 0.060\ g/cm^3$ of material and the other test tube had $0.30\ g/20.0\ cm^3 = 0.015\ g/cm^3$.

The mass of solid dissolved per unit volume of liquid is called the "concentration" of the solution. This unit is the same as that of density, g/cm^3. However, in the case of density, the mass and the volume refer to the same substance. In the case of concentration, the mass refers to the dissolved solid (called the "solute") and the volume refers to the liquid (called the "solvent").

To avoid confusion and for other reasons of convenience, concentrations are often given in $g/100\ cm^3$.

A concentration of $0.015\ g/cm^3$ means that 0.015 g of solute is dissolved in 1 cm^3 of water. A volume of 100 cm^3 of water will, therefore, contain $100 \times 0.015\ g = 1.5\ g$ of solute. Therefore, the concentration of the solution in $g/100\ cm^3$ is $1.5\ g/100\ cm^3$. In general, to find the concentration of a solution in $g/100\ cm^3$, we multiply the concentration in g/cm^3 by 100.

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1 For Experiment 4.1, calculate the concentration of the solutions in g/cm^3 and in $g/100\ cm^3$ after the addition of each sample of solid.

2 What can you say about the largest concentration you were able to make in the experiment?

3 A mass of 25.0 g of sugar is dissolved in 150 cm^3 of water. What is the concentration in $g/100\ cm^3$?

4.3 EXPERIMENT COMPARING THE CONCENTRATIONS OF SATURATED SOLUTIONS

From the results of dissolving the orange solid (Experiment 4.1), you know that you will reach a point where no more solute will dissolve in the solvent. The solution then has the largest possible concentration and is called a saturated solution, as said in section 4.1.

To find the concentration of a saturated solution, you could add a tiny amount of solid at a time and see whether it dissolves. A better method is to begin with a large mass of solid and shake it until you judge that no more will dissolve. Then you can pour off some of the clear liquid and find the concentration.

Try dissolving 5 g of two solids in separate test tubes, each containing 5 cm^3 of water. Stopper the test tubes, and shake them vigorously for several minutes until you have a saturated solution. If the tube cools during the process, keep it warm with your hand.

- Does one sample of solid appear to be more soluble in water than the other?

To find the concentrations of the two saturated solutions, you can evaporate the liquid from a known mass of each solution. Subtracting the mass of the remaining dry solid from the mass of the solution will give you the mass and, therefore, the volume of the water of your sample. This will give you the data you need to calculate the concentration of the saturated solution. You can do the experiment for one solution while some of your classmates work with the other solution.

Pour almost all the saturated solution into a previously massed evaporating dish, being careful not to pour out so much solution that undissolved solid is carried over from the test tube into the dish.

After finding the total mass of dish and solution, you can slowly evaporate the saturated solution to dryness over a flame, as shown in Figure 4.1 (below), and find the mass of the remaining solid. Be careful to heat the solution very slowly so that solid does not spatter out of the dish. Keep watching the dish, and move the flame away whenever spattering begins.

- What was the mass of the solid and the mass of the water in which it dissolved?
- What was the volume of the water?
- What was the concentration of each of the saturated solutions?
- How do your results compare with those of your classmates?

The concentration of a saturated solution is called the "solubility." In this experiment, you found the solubilities of two substances in the same solvent, namely water.

Solubility is independent of the mass of the sample from which it is found. It is a characteristic property of the combination of the solute and the solvent.

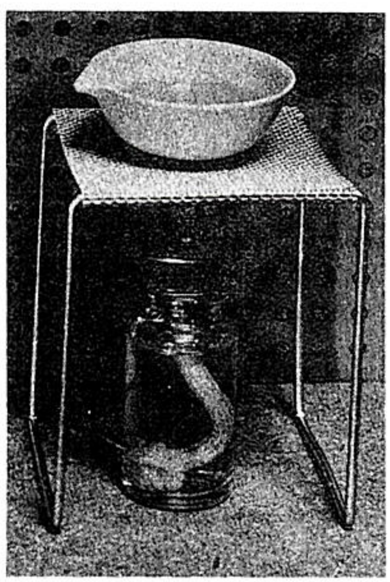

Figure 4.1
Evaporating a solution in an evaporating dish heated over an alcohol burner. If the liquid spatters, it should be heated more slowly by moving the burner to one side so that the flame heats only one edge of the dish.

4 Mario wishes to construct a table that lists the solubility in water of several substances. From various sources he finds the following data for solubilities at 0°C.

a) Boric acid	0.20 g in 10 cm^3 of water
b) Bromine	25 g in 600 cm^3 of water
c) Washing soda	220 g in 1,000 cm^3 of water
d) Baking soda	24 g in 350 cm^3 of water

What is the solubility of each substance in grams per 100 cm^3 of water?

5 From your answers to problem 4, find the largest mass of each substance that will dissolve in 60 cm^3 of water.

6 Suppose that 200 cm^3 of a saturated solution of potassium nitrate were left standing in an open beaker on your laboratory desk for three weeks. During this time most of the water evaporated. Would the mass of potassium nitrate dissolved in the solution change? Would the concentration of the potassium nitrate solution change during the three weeks?

4.4 EXPERIMENT THE EFFECT OF TEMPERATURE ON SOLUBILITY

In the last experiment, you tried to keep the temperature of the solution constant (by warming the test tube with your hand if it cooled). How will the solubility of different substances be affected by the temperature of the liquid? Remember that solubility is the maximum mass of a solid that will dissolve in a given volume of liquid.

To find out, add 10 g of two solids to test tubes, each containing 10 cm^3 of water. Place both test tubes in a large beaker of water, and stir the solutions for several minutes until they are saturated. Now heat the beaker, stirring both solutions constantly, until the water in the beaker is near boiling.

- What do you observe?
- Do the solubilities of the substances appear to change equally or differently as the temperature of the water is increased?

- What do you predict will happen if you remove the burner and cool both test tubes together in a beaker of cold water? Try it.

Figure 4.2 shows the result of an experiment with potassium sulfate. The solubilities at different temperatures were measured by the same method you used in Experiment 4.3. The solubility is expressed as the mass in grams of the substance that is dissolved in 100 cm³ of water to give a saturated solution or, to put it another way, the maximum mass of the substance that can be dissolved in 100 cm³ of water.

Suppose we dissolve 20 g of potassium sulfate in 100 cm³ of water at 80°C. If we now cool the solution to room temperature, 25°C, Figure 4.2 shows that the water can hold only 12 g of potassium sulfate in solution at this temperature. Therefore, during the cooling process, small crystals of solid potassium sulfate begin to appear in the solution at about 70°C. As cooling continues, more crystals are produced in the solution, and they sink to the bottom. A solid that crystallizes out of a saturated solution in this manner is

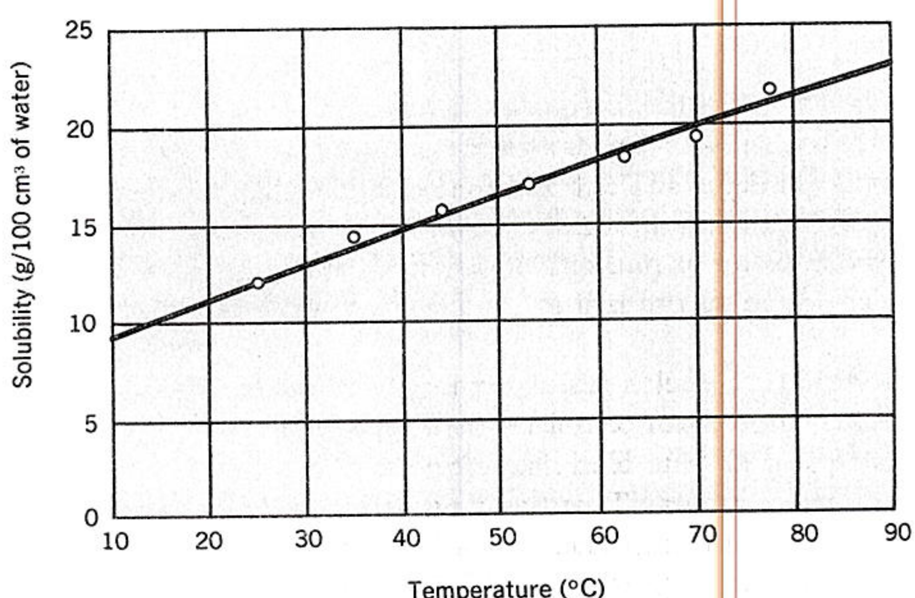

Figure 4.2
A graph of the solubility of potassium sulfate as a function of temperature. The graph shows the maximum concentration of potassium sulfate solution at different temperatures.

called a "precipitate." As you can see from Figure 4.2, the mass of potassium sulfate that will precipitate out of solution and collect at the bottom in this case will be 20 g − 12 g = 8 g.

Figure 4.3 shows the solubility as a function of temperature for several other common substances, all plotted together in the same graph. These curves clearly show that the way the solubility of a substance changes with temperature is a characteristic property that can help to distinguish between different substances. You can see from the graph, for example, that the solubilities of potassium nitrate and sodium chloride (ordinary table salt) are very nearly the same at room temperature (about 25°C) but are widely different at high and low temperatures.

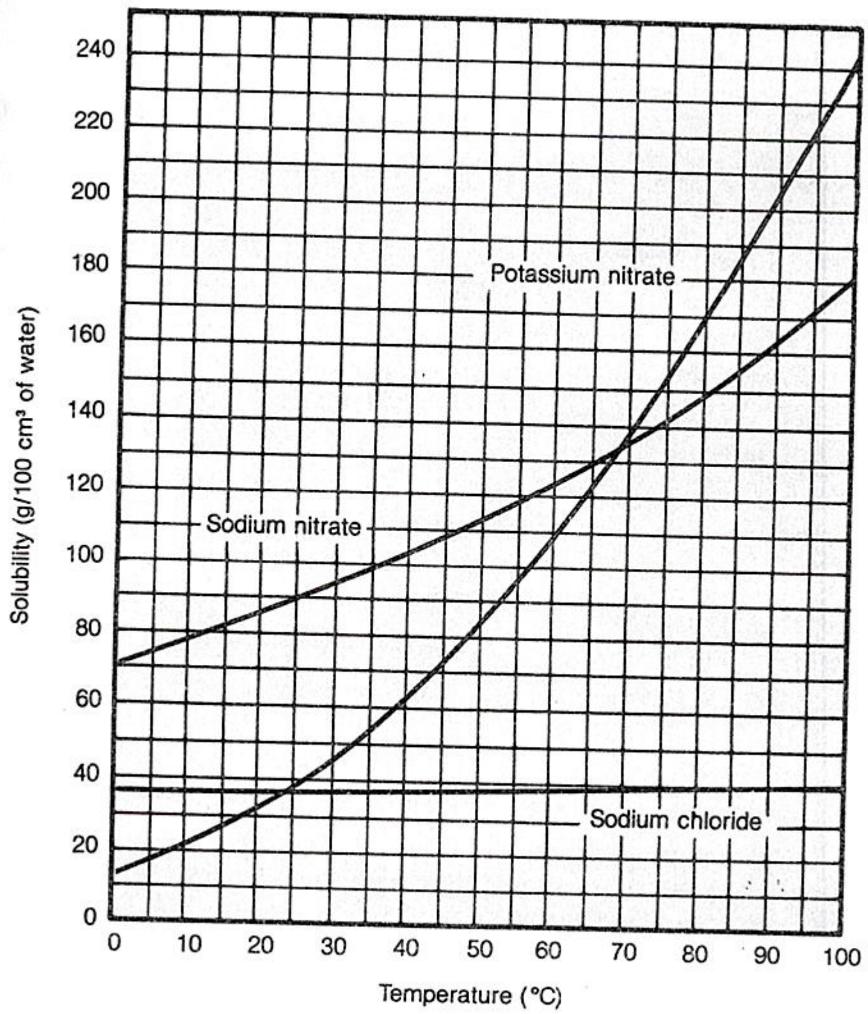

Figure 4.3
Solubility curves of different substances dissolved in water as a function of temperature.

7 If you plotted the data from Figure 4.2 on Figure 4.3, where would the graph be found?

8 Suppose you have a saturated solution of potassium nitrate at room temperature. From Figure 4.3, what do you predict will happen if you

a) heat the solution?

b) cool the solution?

9 What reason can you give for heating the solutions in Experiment 4.4 by immersing the test tubes in a beaker of hot water rather than heating the test tubes directly over a burner flame?

10† What temperature is required to dissolve 110 g of sodium nitrate in 100 cm^3 of water?

11 a) If 20 g of sodium chloride is dissolved in 100 cm^3 of water at 20°C, is the solution saturated?

b) How do you know if a solution is saturated?

12† A mass of 30 g of potassium nitrate is dissolved in 100 cm^3 of water at 20°C. The solution is heated to 100°C. How many more grams of potassium nitrate must be added to saturate the solution?

13 A mass of 10 g of sodium nitrate is dissolved in 10 cm^3 of water at 80°C. As the solution is cooled, at what temperature should a precipitate first appear?

4.5 WOOD ALCOHOL AND GRAIN ALCOHOL

Most rocks and metals, and many other materials, are so slightly soluble in water that we cannot measure the very small amounts that dissolve.

Water, however, is not the only liquid. Perhaps some substances that hardly dissolve at all in water will dissolve easily in other liquids. If such is the case, we can use the different solubilities of substances in these liquids to distinguish between them, as we did with materials soluble in water. We shall first investigate wood alcohol and grain alcohol, two common liquids. Then we shall see if there are other solvents that will further increase our stock of tools for investigating matter.

Wood alcohol, as its name implies, was first made from wood. In fact, some of the liquid that you collected in your distillation of wood was wood alcohol. The ancient Syrians heated wood in order to obtain the liquids and tars that resulted. The watery liquids (including the alcohol mixed with other liquids) were used as solvents and as fuel for lamps. The tars were used to fill the seams in boats, to preserve wood against rot, and as mortar for bricks.

The method used by the Syrians in making these substances was the same as the one you used when you distilled wood, although the apparatus they used was cruder. Lengths of wood were stacked closely in a dishlike depression in the top of a mound of earth. A drain ran from the middle of the depression to a collection pit. After the wood was covered with green branches and wet leaves, a fire was started inside the pile. As this fire smoldered, watery liquids and tars drained off from the pile and collected in the pit. Later it was discovered that the watery liquids could be separated, just as you separated them after you distilled wood. One of these liquids was wood alcohol.

Grain alcohol can be made by fermenting grains, such as corn, barley, and rye, and also by fermenting grapes and other fruits. Fermentation is the process that goes on naturally when fruit juices or damp grain are stored with little exposure to air. Gas bubbles out of the liquid, and what remains boils at a temperature lower than the boiling point of water. As had been discovered long before the beginning of recorded history, this liquid contained a new substance different from water. It was used as a beverage (with effects quite different from those of water) and as a medicine.

Some time before the twelfth century, the wine from fermented grapes was first distilled, and the condensed liquid was described as the "water that burns." It was named *alcohol vini* or "essence of wine," and later came to be called "grain alcohol."

If we measure the densities of the alcohols we get from different grains and fruits, we find no difference between them. These alcohols also have the same boiling point and melting (or freezing) point. In fact, they are all the same substance. Similarly, the alcohols (or essences) produced from different kinds of wood are all

Table 4.1 Some Characteristic Properties of the Most Common Alcohols

	Density (g/cm^3)	Melting point (°C)	Boiling point (°C)
Wood alcohol (methanol)	0.79	−98	64.7
Grain alcohol (ethanol)	0.79	−117	78.5

the same: all have the same density, boiling point, and melting point. Since the distilled liquids obtained from wood and fermented grain are both called alcohol, you might think that they are the same substance. But from an examination of Table 4.1, we see that they are indeed different. Though their densities are nearly the same, their melting points and boiling points differ enough so that there can be no possibility that they are the same substance. Today, wood alcohol is called "methanol," and grain alcohol is called "ethanol."

Both alcohols can be used as fuels. The fuel in your laboratory burner should contain mostly ethanol. Ethanol is the important ingredient in alcoholic beverages; methanol, on the other hand, is very poisonous. Some methanol, or other poison, is usually added to burner fuel to make it undrinkable; ethanol treated this way is said to be "denatured." Both will dissolve many substances that are insoluble in water, and these alcohols have been used as solvents for centuries.

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14 You distilled wood in Experiment 1.1. Where would you expect to find the methanol at the end of the stage of the experiment shown in Figure 1.2? In Figure 1.3?

4.6 EXPERIMENT METHANOL AS A SOLVENT

Sugar and citric acid look the same. They are both white. Sugar has a density of 1.59 g/cm^3, and citric acid has a density of 1.54 g/cm^3. These densities differ slightly—by only about 3 percent. You would find it difficult indeed to determine the volume of an irregular piece of either material to within 3 percent. The values that you would get for the densities of these substances could be no more accurate than your values for the volumes; and if either of these was in error by 3 percent or more, you could not be sure from your density determination whether the two materials were the same substance or not. It is also difficult to distinguish between these substances by their solubilities in water, because both are about equally soluble in water.

Suggest a way to distinguish between these substances by their solubility in methanol, and have your teacher approve it before you start. **CAUTION:** Do not inhale methanol vapor.

- Do the solubilities of sugar and citric acid in methanol help in distinguishing between the two substances?

You can also test the solubilities of other substances, such as moth flakes, magnesium, and magnesium carbonate.

- Do moth flakes dissolve in water? In methanol?
- Do magnesium and magnesium carbonate dissolve in water? In methanol?

You can now see how much we have enlarged our collection of tools for distinguishing between substances. We do not always have to measure density, melting point, or boiling point. For example, suppose we have two test tubes of colorless liquids and we place a piece of copper in each. If the copper dissolves in one but not in the other, we know the two liquids are different. Similarly, if we have two samples of white crystals that dissolve equally well in one solvent, but do not dissolve equally well in another solvent, then they are different substances. You saw an example of this in the case of sugar and citric acid.

15† Two solids appear to be the same and are both insoluble in methanol. A student, whose results are reliable to 5 percent, reports the solubilities of the two solids in water, as in the following table. Are these solids the same substance? Explain your answer.

Solid	Solubility ($g/100\ cm^3$) 0°C	Solubility ($g/100\ cm^3$) 100°C
A	73	180
B	76	230

16 a) Which of the following substances, X, Y, and Z, do you think are the same?

b) How might you test them further to make sure?

Substance	Density (g/cm^3)	Melting point (°C)	Boiling point (°C)	Solubility in water at 20°C ($g/100\ cm^3$)	Solubility in methanol at 20°C
X	1.63	80	327	20	insoluble
Y	1.63	81	326	19	insoluble
Z	1.62	60	310	156	insoluble

4.7 SULFURIC ACID

Another solvent useful in distinguishing between different substances was first produced more than a thousand years ago by heating a soft rock called "green vitriol." The vapors produced by heating the rock were cooled and condensed to form an oily liquid, which was called "oil of vitriol." Its modern name is "sulfuric acid." The ancient method of producing this acid has been abandoned in favor of a far more effective process of manufacture that starts with sulfur. As a result, the acid is readily available in large quantities. It is of great importance in industry and is used in the manufacture of many other substances.

One of the useful properties of sulfuric acid is its ability to dissolve some substances that will not dissolve in water. In many cases, when a substance is dissolved in sulfuric acid, a gas is given off. We shall investigate two such gases in the next experiment.

4.8 EXPERIMENT TWO GASES

One of the substances that produce a gas when dissolved in sulfuric acid is magnesium. To collect the gas, you can use the same apparatus as in Experiment 3.10. (**CAUTION:** Sulfuric acid is highly corrosive, so try not to spill it or get any on your hands, clothing, or books. If you do, wash it off immediately with water and tell your teacher. Be sure to wear your safety glasses.) Use about half a test tube of acid and five 7-cm lengths of magnesium ribbon to produce several test tubes of gas.

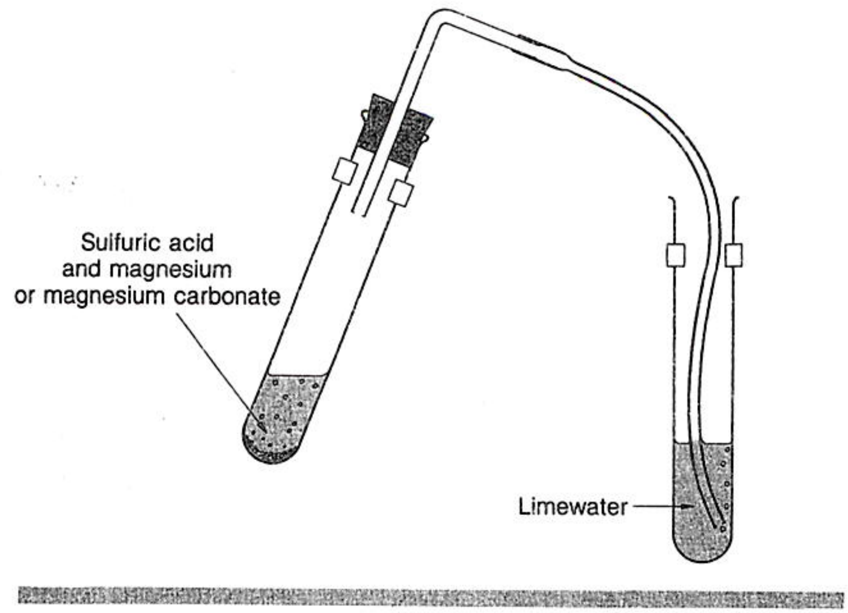

Figure 4.4
The limewater test. Gas from the test tube on the left is bubbled through limewater in the one on the right.

- Should you discard the gas collected in the first test tube? Why?

Try lighting tubes of the gas, using first a burning splint and then a glowing one. Try it holding some of the tubes upside down and some right-side up.

- Does the gas burn?
- Is the gas more dense or less dense than air?

Try bubbling a little of the gas into limewater (Figure 4.4).

Repeat the experiment, using magnesium carbonate instead of magnesium.

4.9 HYDROGEN

The gas you made by dissolving magnesium in sulfuric acid was first produced many centuries ago by the action of sulfuric acid on metals. It was called "inflammable air" because it burned. Later, it was discovered that when inflammable air burns, it produces another gas, which can easily be condensed to a liquid. This liquid was found to boil at 100°C, freeze at 0°C, have a density of 1.00 g/cm^3, and dissolve the same substances that water does. In fact, all the properties of this liquid were the same as those of water, and so it was considered highly unlikely that it could be anything else. Because it produced water when it burned, the gas was given the name "hydrogen" about 200 years ago. "Hydro" is a prefix that means "water"; "gen" is a suffix that means "producing" or "generating."

However, other gases also give water when burned and, like hydrogen, are less dense than air. For many years, these gases were thought to be the same as hydrogen. Marsh gas, now called "methane," is an example that has been known for thousands of years. It is often produced by vegetable matter decaying at the bottoms of lakes or marshes, where it slowly bubbles up to the surface of the water. But this gas does not have nearly as low a density as hydrogen, and so it must be a different substance. It was not until the density of gases could be measured with reasonable accuracy that marsh gas was found to be a different substance from hydrogen.

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17 A student might try to find the density of hydrogen by the method you used for finding the density of a gas in Experiment 3.10, where you collected about 400 cm^3 of gas.

a) What would be the mass of this volume of hydrogen?

b) What difference in mass would there be between the total mass of test tube, acid, and magnesium at the start, and the total mass of test tube and contents at the end of the action?

c) Could you accurately measure this difference in mass on your balance?

4.10 CARBON DIOXIDE

The gas that you produced by dissolving magnesium carbonate in sulfuric acid is called "carbon dioxide." It is a part of the air we exhale at every breath and is also produced by a variety of substances when they are placed in sulfuric acid. The density of carbon dioxide is 1.8×10^{-3} g/cm^3, which is much greater than that of air. Since carbon dioxide does not burn, it makes a good fire extinguisher. It simply smothers the fire like a blanket.

Solid carbon dioxide, called "dry ice," evaporates directly into carbon dioxide gas without first turning into a liquid. But, like melting ice, it stays cold until all the solid is gone. Ice melts at 0°C, but dry ice evaporates at a much lower temperature (at −78.5°C). It remains at this temperature as it evaporates, and for this reason it is often used to cool things down to very low temperatures.

4.11 THE SOLUBILITY OF GASES

If the gases we have investigated so far had been very soluble in water, we could not conveniently have collected them over water. They would simply have dissolved as fast as we produced them. In fact, these gases are so insoluble that it would be difficult to measure the amount of each that dissolves and then distinguish between them by their solubility.

Nevertheless, no gases are completely insoluble in water, and it is fortunate that gases do dissolve. Fish, for example, obtain the oxygen they need from oxygen dissolved in the water around them.

When carbon dioxide is dissolved in water, it changes the properties of water as a solvent. For example, calcium carbonate (the main ingredient of limestone and chalk) is much more soluble in water that contains carbon dioxide than in water that does not. Raindrops dissolve carbon dioxide as they pass through the atmosphere. When rain water enters the soil and hits a layer of limestone, it slowly dissolves the limestone, leaving an empty space. When the soil above this empty space can no longer support itself, it collapses (Figure 4.5).

18† In certain shallow parts of Long Island Sound, fish have been found dead of oxygen starvation on extremely hot days, though at other times fish live there very happily. What property of oxygen do these facts suggest?

19 Experiment 3.10 was performed twice, but with one tablet and 15 cm^3 of water in each case. When the rubber tube was placed as shown in Figure 3.9, 435 cm^3 of gas was collected. When the tube reached only slightly beyond the mouth of the bottle, only 370 cm^3 of gas was collected. The change in mass of the reactants was the same in both cases.

a) What volume of gas dissolved in the water?

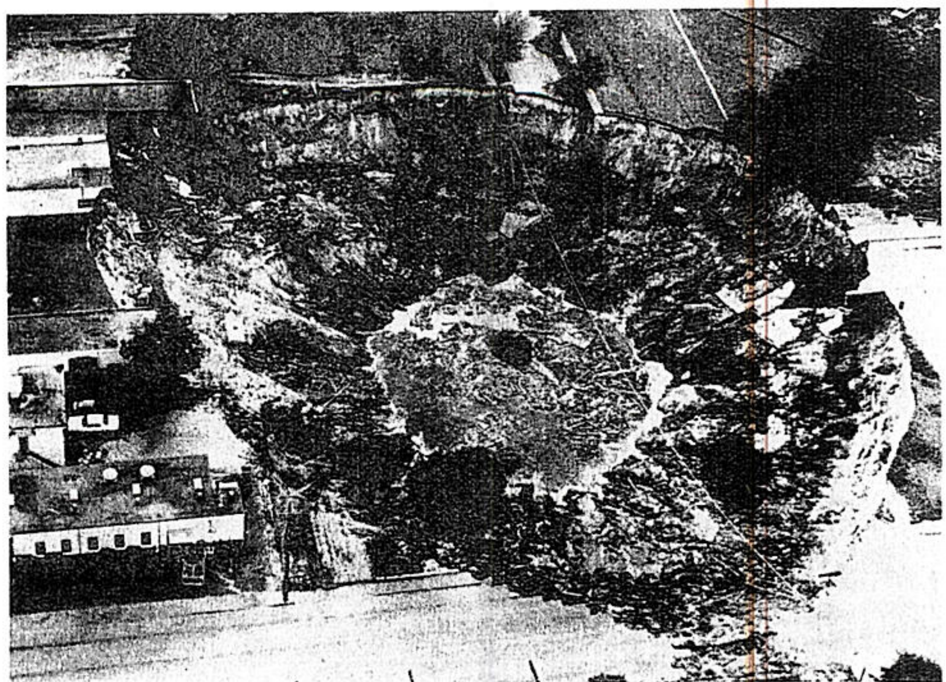

Figure 4.5
A "sinkhole" in central Florida. Limestone below the top layer was dissolved in water containing much carbon dioxide. (*AP/Wide World Photos*)

b) Use Table 3.1 (page 63) to find what mass of gas dissolved in the water.

20 The solubility of chalk in water is 10^{-3} g/100 cm^3. How much water would be necessary to dissolve a piece whose mass is 5 g?

4.12 EXPERIMENT THE SOLUBILITY OF AMMONIA GAS

There are gases that are very soluble in water. A common example is ammonia, whose solution in water is used as a household cleaner. Figure 4.6 shows how ammonia gas can be produced by slowly heating a water solution of ammonia. The gas can then be collected in a dry test tube.

Collect a test tube of the gas. (Don't try to smell it directly—it is very irritating.) Then, very slowly, remove the tube and close it

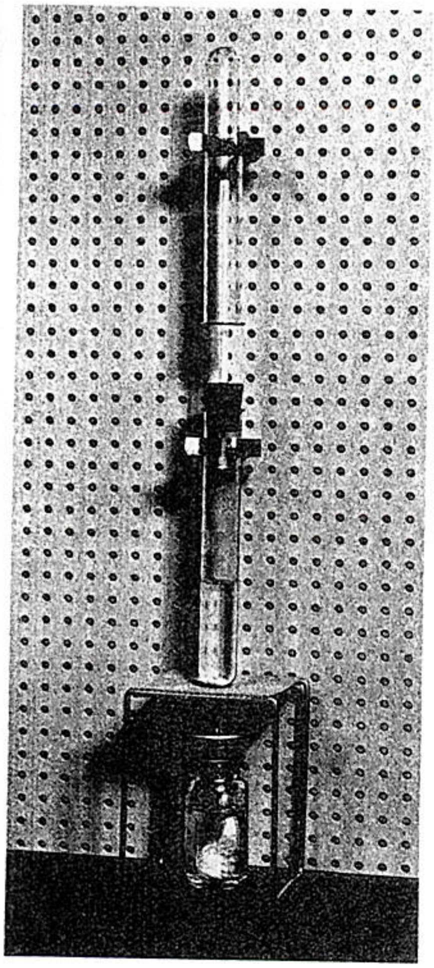

Figure 4.6
Collecting ammonia gas produced by heating a water solution of the gas.

with a stopper. Place it, mouth downward, in a beaker of water, and then remove the stopper.

- What do you observe?
- From the way you produced the gas, what can you conclude about the solubility of the gas as the temperature of the solution is raised?
- From the method of collecting the gas, what can you conclude about its density?

21† If the tube of ammonia gas had been inverted and placed in water saturated with ammonia, would the liquid have risen in the test tube as it did in this experiment?

For Home, Desk, and Lab

22 Each of four test tubes contains 10 cm^3 of water at 25°C. The following masses of an unknown solid are placed in the test tubes: 4 g in the first, 8 g in the second, 12 g in the third, and 16 g in the fourth. After the tubes are shaken, it is observed that all of the solid has dissolved in the first two tubes, but that there is undissolved solid in the remaining two tubes.

a) What is the concentration of the solid in each of the first two tubes?

b) What can you say about the concentration of the solid in the second two tubes?

23 a) Which of the substances shown in Figure 4.3 could be the unknown solid of problem 22?

b) If the unknown is indeed the substance you named in answer to part (a), what will happen if the solution in each test tube is cooled to 10°C?

24 The solubility of a substance in water was found to be 5 g/100 cm^3 at 25°C, 10 g/100 cm^3 at 50°C, and 15 g/100 cm^3 at 75°C. What would you expect its solubility to be at 60°C? At 100°C? Explain.

25 In many localities, after a kettle has been used for some time for boiling water, a flaky solid appears on the inside-bottom and on the sides of the kettle that have been in contact with the water. How do you account for the presence of this "boiler scale"?

26 Your experiment with moth flakes (Experiment 4.6) showed that this substance was insoluble in water but that it dissolved readily in methanol.

a) Predict the effect of adding water to a methanol solution of moth flakes. Try it.

b) Sugar proved to be almost insoluble in methanol but dissolved readily in water. Predict the effect of adding methanol to a solution of sugar in water. Try this too.

27 There are two kinds of felt-tip (magic-marker) pens. Some are labeled "permanent" and some are labeled "water color."

a) What does the label tell you about the solubility in water of the dye in the two inks?

b) Do you think the liquid in both inks is water?

28 In dry cleaning, a garment is sprayed with liquids that dissolve various stains. Often a brightly colored cotton shirt carries a label: "Dry clean only—colors may bleed when washed."

a) What does this tell you about the solubility of the dye in hot water containing a detergent or soap?

b) Do you think the dye dissolves in cold tap water?

29 In Experiment 4.8, you found that magnesium metal would dissolve in sulfuric acid.

a) Does this observation enable you to predict with certainty that all metals will dissolve in sulfuric acid?

b) Try dissolving other metals, such as copper, zinc, lead, and aluminum, in sulfuric acid.

30 When magnesium carbonate is placed in sulfuric acid, a gas is produced whose properties you studied in Experiment 4.8. When you place washing soda in hydrochloric acid, you also get a gas. What would you do with this gas to determine whether it is the same gas that you got from dissolving magnesium carbonate in sulfuric acid?

31 Three samples of gas are tested for characteristic properties. Sample A does not turn limewater milky, is less dense than air, and burns. Sample B turns limewater milky, is more dense than air, and burns. Sample C turns limewater milky, is less dense than air, and burns. What can you conclude about these samples of gas?

32 A fizzing tablet is dissolved in 10 cm^3 of water, and the gas is collected as in Experiment 3.10. The volume of gas collected is 450 cm^3. When 50 cm^3 of water is used, the volume of gas collected is 405 cm^3. The tube was all the way up in the bottle in both cases.

a) Why do you think the volume is less?

b) Would this make a difference in the density calculation?

33 The following table shows the solubility of carbon dioxide at atmospheric pressure at various temperatures.

Solubility ($g/100\ cm^3$)	Temperature (°C)
0.34	0
0.24	10
0.18	20
0.14	30
0.12	40
0.10	50
0.086	60

a) Draw a graph from these data.

b) Find, from your graph, how much carbon dioxide will dissolve in 100 cm^3 of water at 25°C.

c) Use Table 3.1, page 63, to find the volume of carbon dioxide that will dissolve in 100 cm^3 of water at room temperature (20°C).

d) How much carbon dioxide will dissolve in 100 cm^3 of water at 95°C?

34 If you have a certain amount of a solid to dissolve in water, you usually can hasten the process in various ways. Why do you think each of the following steps is effective in making the solid dissolve faster?

a) Stirring the water
b) Crushing the solid into smaller particles
c) Heating the water

Theme for a Short Essay

Write an episode in a mystery centered on the possibility of selectively dissolving some substances mixed in with others. For example, a secret message written in ink may be covered by a painting. Describe real (not fictitious) substances.