Properties of Solutions

TOPIC 7

What You Know About Solutions

What can you tell about the saturation of a solution using a crystal of the solute?

Can you do an experiment to determine if a solution is saturated?

Sure. Take a crystal of the solute and drop it into the solution, then observe what happens:

If it dissolves, the solution is unsaturated.

If it falls to the bottom, the solution is saturated.

If more crystals form, the solution is supersaturated.

Learn more about solutions and their properties while studying this topic.

TOPIC 7

Properties of Solutions

Vocabulary

boiling point	percent mass	supersaturated
ion-dipole forces	saturated	unsaturated
molarity	solute	vapor
parts per million (ppm)	solution	vapor pressure
percent by volume	solvent	

Topic Overview

Most of the materials that you use every day are not pure substances. It is more likely that they are mixtures. This topic will explore an important type of mixture, the solution. The nature and properties of solutions are important concepts used in chemistry. One reason they are so important is that most chemical reactions take place in solutions. In this topic you will study the nature and properties of solutions and ways to express the concentration of solutions.

Solutions

A **solution** is a homogeneous mixture of substances in the same physical state. Solutions contain atoms, ions, or molecules of one substance spread uniformly throughout a second substance. When salt (NaCl) is stirred into water, the individual ions of the salt separate and uniformly spread throughout the water, forming a solution.

Types of Solutions

A solid may be dissolved in another solid. Brass is a mixture of zinc and copper. When metals are mixed to form a solution, the result is called an alloy. Air is an example of a mixture of gases forming a solution.

Memory Jogger

Mixtures do not have definite composition. For example, air is a mixture. The percentage of water vapor present in the air varies from day to day. Photosynthesis increases the concentration of oxygen in the air and reduces carbon dioxide. Respiration has the opposite effect, decreasing oxygen and increasing carbon dioxide.

Although solutions exist in all three states, the discussion in this topic will be limited to liquid solutions. Perhaps the most common type of solution is one in which a solid or a liquid is dissolved in a liquid.

The terms *solute* and *solvent* are commonly used to identify the parts of a solution. In general terms, the **solute** is the substance that is being dissolved, and it is the substance present in the smaller amount. When solid sodium nitrate dissolves in water, the sodium nitrate is the solute. The substance that dissolves the solute is the **solvent**, and it is present in the greater amount. Water is, perhaps, the most common solvent. Water solutions are called aqueous solutions, and the notation (aq) is used in equations to show that the substance is dissolved in water.

$$NaCl(s) \rightarrow Na^{+}(aq) + Cl^{-}(aq)$$

Once the salt and water are stirred and the mixture becomes homogeneous, the dissolved particles will not settle. Liquid solutions are clear, and light will pass through a solution without being dispersed, as shown in Figure 7-1.

Solutions may or may not have color. For example, solutions of copper salts have a characteristic blue color, while a solution of sodium nitrate is colorless.

The following list summarizes characteristics of liquid solutions.

- Solutions are homogeneous mixtures.
- Solutions are clear and do not disperse light.
- Solutions can have color.
- Solutions will not settle on standing.
- Solutions will pass through a filter.

Memory Jogger

Dissolved particles are small enough that they will pass through a filter, so filtration cannot be used to separate the parts of a solution. Distillation is one method that can be used to separate the components of a solution. If a solution consists of a solid dissolved in a liquid, distillation removes the solvent, leaving the solute behind.

Solubility Factors

You've noticed that some things easily dissolve in water or other solvents. When you make a cup of coffee, certain materials in the coffee grounds dissolve but other materials don't. Sugar will readily dissolve in the cup of coffee but the spoon you use to stir the solution does not dissolve. How much of a solute will dissolve in a certain amount of solvent at a certain temperature is known as solubility. Materials with a high solubility are said to be soluble; materials with a low solubility are said to be insoluble. What factors determine the solubility of a solute in a solvent?

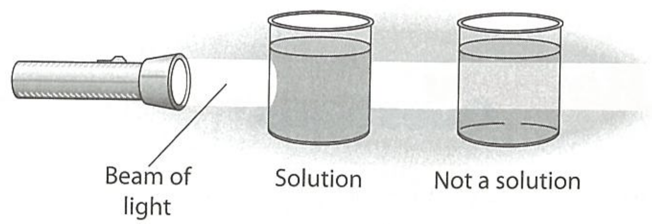


Figure 7-1. Light passing through a solution: The particles in a solution are too small to disperse light. When light passes through a liquid with larger particles, such as gelatin in water, you can see the beam because larger particles disperse light.

Nature of Solute and Solvent When sodium chloride dissolves in water it does so because its positively and negatively charged ions are attracted to the oppositely charged ends of the polar water molecule. The dissolving process is shown in Figure 7-2. The positively charged sodium ions are attracted to the negative pole of the water molecules. In like manner, the negatively charged chloride ions are attracted to the positive end of the water dipole and are dissolved. The attractive forces between the ions and water molecules are called **ion-dipole forces**, and are greater than the forces of attraction between the ions themselves. Ionic and polar substances dissolve in polar solvents.

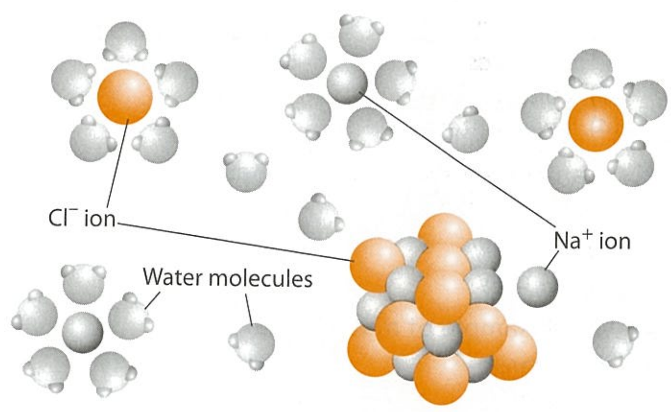


Figure 7-2. The dissolving process: Ionic and polar solutes dissolve in polar solvents because unlike charges attract each other.

Nonpolar substances, such as fats, do not dissolve in water because there aren't strong attractive forces between the fat molecules and the water molecules. Fat molecules will dissolve in nonpolar solvents. The forces that hold the nonpolar molecules to each other are quite weak, and the molecules simply mix together. The term "like

Table 7-1. Solubility Summary

Solute Type	Nonpolar Solvent	Polar Solvent
nonpolar	soluble	insoluble
polar	insoluble	soluble
ionic	insoluble	soluble

dissolves like" is often used to describe what solutes will dissolve in what solvents. Table 7-1 summarizes this concept.

An interesting and important case of "like dissolves like" is found in the action of soaps. Greases, which are nonpolar, won't easily wash off our hands in water, which is polar. Soaps are long carbon chains that have one end that is polar, allowing the soap to dissolve in water. The other end of the soap is nonpolar, and grease will dissolve in it.

Temperature As temperature increases, most solids become more soluble in water. A few exceptions exist. Gases react in the opposite manner. As temperature rises, the solubility of all gases in liquids decreases.

Pressure Pressure has little or no effect on the solubility of solid or liquid solutes. Pressure does affect the solubility of gases in liquids. As pressure increases, the solubility of gases in liquids increases. When a can of soda is opened, the pressure decreases. The carbon dioxide is no longer as soluble at the lowered pressure, and it escapes as bubbles.

Review Questions — Set 7.1

1. In a true solution, the dissolved particles
(1) are visible to the eye
(2) will settle out on standing
(3) are always solids
(4) cannot be removed by filtration

2. When a teaspoon of sugar is added to water and stirred, the sugar
(1) melts
(2) dissolves
(3) condenses
(4) evaporates

3. In an aqueous solution of potassium chloride, the solute is
(1) Cl^- only
(2) K^+ only
(3) K^+Cl^-
(4) H_2O

4. Which sample of matter is a mixture?
(1) $H_2O(s)$
(2) $H_2O(\ell)$
(3) $NaCl(\ell)$
(4) $NaCl(aq)$

5. Most ionic substances are soluble in water because water molecules are
(1) nonpolar
(2) inorganic
(3) ionic
(4) polar

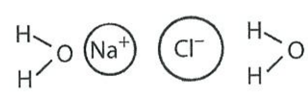
6. An aqueous solution of copper sulfate is poured into a filter paper cone. What passes through the filter paper?
(1) only the solvent
(2) only the solute
(3) both solvent and solute
(4) neither the solute nor solvent

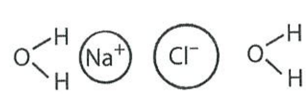
7. Nonpolar solvents will most easily dissolve solids that are
(1) ionic
(2) covalent
(3) metallic
(4) colored

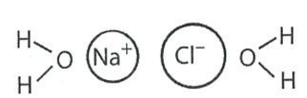
8. As the temperature rises, the solubility of all gases in water
(1) decreases
(2) increases
(3) remains the same

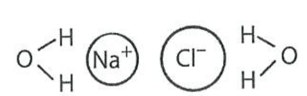
9. A decrease in pressure has the greatest effect on a solution that contains
(1) a gas in a liquid
(2) a liquid in a liquid
(3) a solid in a solid
(4) a solid in a liquid

10. Which diagram best illustrates the ion-molecule attractions that occur when the ions of $NaCl(s)$ are added to water?

(1) 

(2) 

(3) 

(4) 

11. What happens when NaCl(*s*) is dissolved in water?

(1) Cl^- ions are attracted to the oxygen atoms of the water.
(2) Cl^- ions are attracted to the hydrogen atoms of the water.
(3) Na^+ ions are attracted to the hydrogen atoms of the water.
(4) No attractions are involved; the crystal just falls apart.

12. Two grams of potassium chloride are completely dissolved in a sample of water in a beaker. This solution is classified as

(1) an element.
(2) a compound.
(3) a homogeneous mixture.
(4) a heterogeneous mixture.

Looking at Solubility

Solubility information may be presented in different ways. Solubility Curves in the *Reference Tables for Physical Setting/Chemistry* presents quantitative information showing the relationship of grams of solute that may be dissolved at various temperatures. Solubility Guidelines provides some general guidelines about the solubility of ionic substances. You will need to be able to interpret information from both tables.

Solubility Graphs The Solubility Curves table shows the number of grams of a substance that can be dissolved in 100. g of water at temperatures between 0°C and 100°C. Each line represents the maximum amount of that substance that can be dissolved at a given temperature. All of the lines that show an increase in solubility as temperatures increase represent solids being dissolved in water. Although these lines on the graph show an increase in solubility as temperature increases, a few solids, such as cesium sulfate, become less soluble as temperature increases.

Three lines show decreasing solubility with increasing temperature. These three lines represent the gases NH_3, HCl, and SO_2. The solubility of all gases decreases with increasing temperature.

Figure 7-3 shows four positions relative to a line of maximum solubility. Position A is below the maximum line of solubility. At this position, the temperature is 35°C, and 25 g of solute X dissolve. Because at this point the solution holds less solute than the maximum it can hold, the solution is said to be **unsaturated.** If there is not a temperature change, an additional 30 g can be added to bring X to position B.

Position B is on the line of maximum solubility. A solution that contains a maximum amount of solute that will dissolve at a specific temperature is **saturated.** At this position, the solution contains 55 g of X. The addition of more solid solute will not result in more being dissolved. Any additional solid that is added will simply settle to the bottom of the container.

If the temperature is reduced to 20°C, only 35 g of X can dissolve. When the temperature of the solution at B is reduced, the most likely event is that the excess 20 g of X will precipitate, and the solution will remain

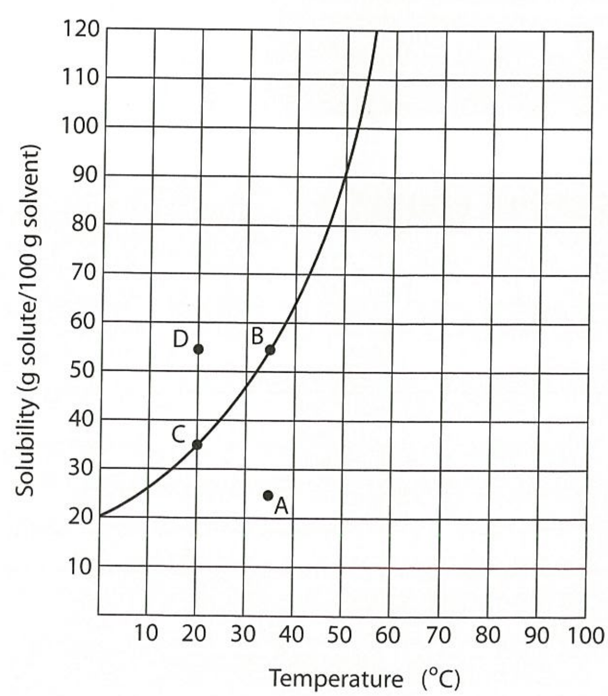


Figure 7-3. A solubility curve

Digging Deeper

There are two terms that are somewhat useful in describing concentrations of solutions. Dilute solutions contain relatively small amounts of dissolved solute in a large amount of solvent. For example, 5 g of a substance dissolved in 100. g of H_2O would be a dilute solution. Concentrated solutions contain relatively large amounts of solute. These terms should not be confused with saturated and unsaturated. Five grams of potassium chlorate dissolved in 100. g of water at 0°C would be dilute, but the solution would be saturated; 80. g of $NaNO_3$ dissolved in the same amount of water would be concentrated, but unsaturated.

Memory Jogger

Recall that double-replacement reactions have the general formula:

AB + CD → AD + CB

saturated at point C. On rather rare occasions, as the temperature decreases, crystals do not form and the substance may be at position D.

At position D, there is more solid dissolved than normal. A solution that holds more solute than is present in a saturated solution at that temperature is **supersaturated.** These solutions are quite unstable. The addition of a single solid crystal of the substance will cause additional solid to form, and the solution will return to a saturated condition. If no temperature change occurred, the solution would become saturated at point C, with 20 g of the substance precipitating. The only way to make a supersaturated solution is to cool a saturated solution in which there are no crystals or impurities, such as dust, present.

Solubility Tables The Solubility Guidelines of *Reference Tables for Physical Setting/Chemistry* contains some guidelines for the solubility of common ionic compounds. The table shows that all compounds of the ammonium and the nitrate ion are soluble. All of the halide ions, such as Cl^-, form compounds that are soluble, but three exceptions are listed. Silver chloride is not soluble, nor are Pb^{2+} nor Hg_2^{2+} chlorides, and they are precipitates if they form in a double-replacement reaction. This table is useful in predicting whether or not a precipitate will form when two ionic solutions are mixed. A reaction will take place if one or both of the products is listed as insoluble.

Recognizing Unsaturated, Saturated, and Supersaturated Solutions Because solutions are clear, it is difficult to simply look at a solution and determine whether it is unsaturated, saturated, or supersaturated.

One method of recognizing the type of solution narrows the choices. If a solution contains some undissolved solute, it must be a saturated solution.

The addition of a solute crystal can also be used to determine its condition. If it dissolves, the original solution was unsaturated. If it simply falls to the bottom, the solution is saturated. If it causes additional crystals to form, the original solution was supersaturated.

SAMPLE PROBLEM

Silver nitrate and sodium chromate solutions are mixed together. Will a precipitate form? If so, what is the name of the precipitate?

SOLUTION: Identify the known and unknown values.

Known	Unknown
Formulas of reactants	Solubility of products
Solubility table	Identity of precipitate

1. Write the word equation for a double-replacement reaction between silver nitrate and sodium carbonate.

silver nitrate + sodium chromate →
silver chromate + sodium nitrate

2. Check the solubilities of the products.

Chromates are listed as insoluble, with the exception of Group 1 ions or the ammonium ion. Nitrates are soluble.

A precipitate of silver chromate will form.

Review Questions

Use the tables in *Reference Tables for Physical Setting/Chemistry* as needed in answering the following questions.

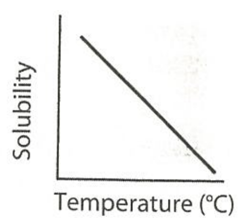
13. Which compound's solubility decreases most rapidly as the temperature changes from 10°C to 70°C?

(1) NH_4Cl (2) NH_3 (3) HCl (4) KCl

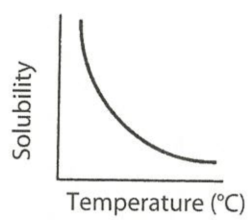
14. Solubility for salt X is shown in the table below.

Temperature (°C)	Solubility (g salt *X*/100 g H_2O)
10	5
20	9
30	13
40	18
50	27
60	35

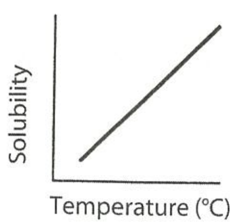
Which graph most closely represents the data shown in the table?



(1)

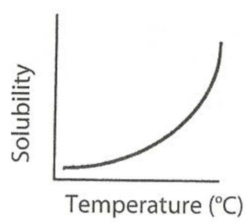


(3)



(2)

(3)



(4)

15. A solution contains 14 g of KCl in 100. g of water at 40°C. What is the minimum amount of KCl that must be added to make this a saturated solution?

(1) 14 g (2) 19 g (3) 25 g (4) 44 g

16. Which salt has the greatest change in solubility between 30°C and 50°C?

(1) KNO_3 (2) KCl (3) $NaNO_3$ (4) NaCl

17. Which of the following substances is least soluble in 100. g of water at 50°C?

(1) NaCl (2) KCl (3) NH_4Cl (4) HCl

18. A student obtained the following data in a chemistry laboratory.

Trial	Temperature (°C)	Solubility (g KNO_3/100 g H_2O)
1	25	40
2	32	50
3	43	70
4	48	60

Based on the reference tables, which of the four trials listed seems to be in error?

(1) 1 (2) 2 (3) 3 (4) 4

19. How many grams of the compound potassium chloride (KCl) must be dissolved in 200. g of water to make a saturated solution at 60°C?

(1) 30 g (2) 45 g (3) 56 g (4) 90 g

20. Which of the following is insoluble?

(1) calcium chloride
(2) ammonium phosphate
(3) barium sulfate
(4) potassium chromate

21. Which of the following is not soluble?

(1) sodium chromate
(2) lead(II) iodide
(3) lithium hydroxide
(4) ammonium sulfate

22. If solutions of sodium chromate and potassium carbonate are mixed and poured into a filter, what will pass through the filter?

(1) sodium, chromate, potassium, and carbonate ions
(2) sodium and carbonate ions only
(3) potassium and chromate ions only
(4) water only

23. Which amount of a compound dissolved in 100. g of water at the stated temperature represents a solution that is saturated?

(1) 20 g $KClO_3$ at 80°C
(2) 40 g KNO_3 at 25°C
(3) 40 g KCl at 60°C
(4) 60 g $NaNO_3$ at 40°C

Concentration of Solutions

Because solutions are homogeneous mixtures, their compositions can vary. Sometimes, it is adequate to refer to a solution as dilute or concentrated. However, *dilute* and *concentrated* are relative terms and are not precise regarding the amount of solute involved. In most cases it is the specific amount, or concentration, of the solute that is important. In this section you will learn several methods of expressing the specific concentration of solute in a solution.

Molarity

One of the most important methods of stating the concentration of a solution is in terms of the number of moles of solute in a given volume of solution. The **molarity** (M) of a solution is the number of moles of solute in 1 L of solution. The relationship is listed in Important Formulas and Equations of *Reference Tables for Physical Setting/Chemistry.*

$$\text{molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

SAMPLE PROBLEM

What is the molarity of a solution that contains 4.0 mol of NaOH in 0.50 L of solution?

SOLUTION: Identify the known and unknown values.

Known	*Unknown*
amount NaOH = 4.0 mol	molarity = ? M
volume of solution = 0.50 L	

Substitute known values into the equation for molarity, and solve for molarity.

$$M = \frac{\text{moles of solute}}{\text{liters of solution}}$$

$$M = \frac{4.0 \text{ moles NaOH}}{0.50 \text{ liter}}$$

$$M = 8.0$$

In the previous sample problem, the molarity of the solution is 8.0 M. A liter of this solution would contain 8.0 mol of the solute, NaOH. However, in many problems, the mass of solute is given instead of the number of moles. To solve this type of problem, convert grams of solute to moles of solute and solve as above.

SAMPLE PROBLEM

What is the molarity of a solution containing 82.0 g of $Ca(NO_3)_2$ in 2.0 liters of solution?

SOLUTION: Identify the known and unknown values.

Known	*Unknown*
mass of $Ca(NO_3)_2$ = 82.0 g	molarity = ? M
volume of solution = 2.0 L	

1. Calculate the formula mass of $Ca(NO_3)_2$.

For Ca:	1 atom × 40. amu/atom	=	40 amu
For N:	2 atoms × 14 amu/atom	=	28 amu
For O:	6 atoms × 16 amu/atom	=	96 amu
	Formula mass $Ca(NO_3)_2$	=	164 amu

2. Change the formula mass to gram formula mass.

gram formula mass = formula mass in grams
gram formula mass $Ca(NO_3)_2$ = 164 g/mol

3. Convert grams of $Ca(NO_3)_2$ to moles.

$$\text{moles } Ca(NO_3)_2 = \frac{\text{grams } Ca(NO_3)_2}{\text{gram formula mass } Ca(NO_3)_2}$$

$$\text{moles } Ca(NO_3)_2 = \frac{82.0\cancel{g}}{164\ \cancel{g}/\text{mol}} = 0.500\ \text{mol}$$

4. Calculate molarity.

$$M = \frac{\text{moles of solute}}{\text{liters of solution}}$$

$$M = \frac{0.500\ \text{mol NaOH}}{2.0\ \text{L solution}}$$

$$M = 0.250$$

Percent by Mass

It is common to find labels that list the concentration of the ingredients by percent mass. Fertilizers often list the active ingredients as a percentage of the entire mass of the fertilizer. **Percent mass** is simply the mass of an ingredient divided by the total mass, expressed as a percent (parts per hundred). Percent mass problems are essentially the same as the percent composition problems found in Topic 3. To calculate the percent mass, use the following relationship.

$$\textbf{percent mass} = \frac{\textbf{mass of part}}{\textbf{mass of whole}} \times \textbf{100\%}$$

Percent by Volume

When two liquids are mixed to form a solution, it is common to express the concentration of the solute as a percent by volume. A label on a bottle of rubbing alcohol shows a common example. Usually the label will show that the solution is 70% isopropyl alcohol by volume. The rest of the solution is water.

Percent by volume is the ratio of the volume of an ingredient divided by the total volume and expressed as a percent.

$$\textbf{percent by volume} = \frac{\textbf{volume of solute}}{\textbf{volume of solution}} = \textbf{100\%}$$

MemoryJogger

In math class you have learned how to rearrange an equation to solve for an unknown variable. Analyze these formulas to remind yourself how to do it.

$$a = bc$$
$$b = a/c$$
$$c = a/b$$

Recognize that the molarity equation can also be rearranged to solve for any variable included in the equation.

$$\text{molarity} = \frac{\text{moles solute}}{\text{liters of solution}}$$

moles solute = (molarity)(liters of solution)

$$\text{liters of solution} = \frac{\text{moles solute}}{\text{molarity}}$$

SAMPLE PROBLEM

What is the percent mass of sodium hydroxide if 2.50 g of NaOH are added to 50.00 g of H_2O?

SOLUTION: Identify the known and unknown values.

Known	*Unknown*
mass of NaOH = 2.50 g	Percent mass NaOH ?%
mass of H_2O = 50.00 g	

Substitute known values into the percent mass equation, and solve for percent mass.

$$\% \text{ mass} = \frac{\text{mass of part}}{\text{mass of whole}} \times 100\ \%$$

$$\% \text{ mass NaOH} = \frac{2.50\ \cancel{g}\ \text{NaOH}}{(2.50 + 50.00)\ \cancel{g}\ \text{solution}} \times 100\%$$

$$\% \text{ mass NaOH} = 4.76\%$$

SAMPLE PROBLEM

What is the percent by volume of alcohol if 50.0 mL of ethanol is diluted with water to form a total volume of 300. mL?

SOLUTION: Identify the known and unknown values.

Known	*Unknown*
volume of ethanol = 50.0 mL	percent by volume ethanol = ? %
total volume = 300. mL	

Substitute known values into the percent by volume equation, and solve for percent by volume.

$$\% \text{ by volume} = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100\%$$

$$\% \text{ by volume ethanol} = \frac{50.0\ \cancel{mL}\ \text{ethanol}}{300.\ \cancel{mL}\ \text{solution}} \times 100\%$$

$$\% \text{ by volume ethanol}\ \ 16.7\%$$

Parts per Million

Parts per million is similar to percent composition because it compares masses. **Parts per million (ppm)** is a ratio between the mass of a solute and the total mass of the solution. This method of reporting concentrations is useful for extremely dilute solutions when molarity and percent mass would be difficult to interpret. For example, chlorine is used as a disinfectant in swimming pools. Only about 2 g of chlorine per 1,000,000 g of swimming pool water is necessary to keep the pool sanitized. Finding molarity and percent mass would result in numbers too small to be useful. Parts per million is often used to report a measured amount of air or water pollutants.

Percent composition uses the amount present per hundred parts because it is a percent. The only difference in finding ppm is that you multiply by 1,000,000 ppm instead of 100 percent.

$$\textbf{ppm} = \frac{\textbf{grams of solute}}{\textbf{grams of solution}} \times \textbf{1,000,000 ppm}$$

SAMPLE PROBLEM

Approximately 0.0043 g of oxygen can be dissolved in 100. mL of water at 20°C. Express this in terms of parts per million.

SOLUTION: Identify the known and unknown values.

Known
mass of O_2 = 0.0043 g
volume of H_2O = 100. mL

Unknown
ppm O_2 = ?

Substitute known values into the ppm equation, and solve for ppm O_2.

$$\text{ppm} = \frac{\text{grams of solute}}{\text{grams of solution}} \times 1{,}000{,}000 \text{ ppm}$$

$$\text{ppm } O_2 = \frac{0.0043 \not{g}}{100.0043 \not{g}} \times 1{,}000{,}000 \text{ ppm}$$

ppm O_2 = 43 ppm

Preparation of a Solution of Known Concentration

It is important to be able to calculate the amount of solute to be added to a known volume of solvent to make a solution of specified concentration. The following sample problem shows you how to determine the amount of solute needed to prepare a solution of known molarity.

You now know how much solute and solvent you need to actually prepare the solution in the sample problem, but the procedure used in preparation

SAMPLE PROBLEM

What mass of sodium carbonate is required to prepare 2.00 L of a 0.250 M sodium carbonate solution?

SOLUTION: Identify the known and unknown values.

Known
concentration of solution = 0.250 M
volume of solution = 2.00 L

Unknown
mass of Na_2CO_3 = ? g

Determine the number of moles of solute needed by using molarity and volume.

moles = MV = molarity × liters of solution

$$\text{moles } Na_2CO_3 = \frac{0.250 \text{ mol}}{1 \not{L}} \times 2.00 \not{L} = 0.500 \text{ mol}$$

The following steps are used to convert moles Na_2CO_3 to grams Na_2CO_3:

1. Determine the formula mass of Na_2CO_3.

For Na: 2 atoms × 23.0 amu/atom = 46.0 amu
For C: 1 atom × 12.0 amu/atom = 12.0 amu
For O: 3 atoms × 16.0 amu/atom = 48.0 amu
formula mass Na_2CO_3 = 106.0 amu

2. Change the formula mass to gram formula mass.

gram formula mass = formula mass in grams
gram formula mass Na_2CO_3 = 106.0 g/mol

3. Convert moles of Na_2CO_3 into grams of Na_2CO_3.

mass = moles × gram formula mass
mass Na_2CO_3 = 0.500 $\not{mol}$ × 106.0 g/$\not{mol}$
mass Na_2CO_3 = 53.0 g

is also essential. The steps listed below apply to the preparation of any solution of known concentration:

1. Add the desired amount of solute to a volumetric flask.
2. Add some distilled water and mix until the solute is dissolved and the solution is homogeneous.
3. Fill the volumetric flask to the mark on the neck of the flask, stopper, and again mix to ensure that the solution is homogeneous.

The reason the water is added in two steps is that it is easier to dissolve the solute if the flask is not full and there is room for the water to be adequately stirred or shaken.

Review Questions — Set 7.3

24. What is the molarity of a KF(aq) solution containing 116 g of KF in 1.00 L of solution?

(1) 1.00 M (2) 2.00 M (3) 3.00 M (4) 4.00 M

25. What is the molarity of an H_2SO_4 solution if 0.25 L of the solution contains 0.75 mol of H_2SO_4?

(1) 0.33 M (2) 0.75 M (3) 3.0 M (4) 6.0 M

26. What is the total number of moles of the solute H_2SO_4 needed to prepare 5.0 L of a 2.0 M solution of H_2SO_4?

(1) 2.5 mol (3) 10. mol
(2) 5.0 mol (4) 20. mol

27. What volume of a 2.0 M solution is needed to provide 0.50 mol of NaOH?

(1) 0.25 L (2) 0.50 L (3) 1.0 L (4) 2.0 L

28. What is the molarity of a solution that contains 40. g of NaOH in 0.50 L of solution?

(1) 1.0 M (2) 2.0 M (3) 0.50 M (4) 0.25 M

29. If 100. mL of a 1.0 M solution is evaporated to a volume of 25 mL, what will be the concentration of the resulting solution?

(1) 0.25 M (2) 0.50 M (3) 2.0 M (4) 4.0 M

30. What is the percent by mass of a solution in which 60. g of NaOH are dissolved in sufficient water to make 100 g of solution?

(1) 16% (2) 40% (3) 60% (4) 160%

31. What is the percent by mass of a solution if 60. g of acetic acid are added to 90. g of water?

(1) 20% (2) 30% (3) 40% (4) 67%

32. Carbon dioxide gas has a solubility of 0.0972 g/100 g H_2O at 40°C. Expressed in parts per million, this concentration is closest to

(1) 0.972 ppm (3) 97.2 ppm
(2) 9.72 ppm (4) 972 ppm

33. A substance has a solubility of 350 ppm. How many grams of the substance are present in 1.0 L of a saturated solution?

(1) 0.0350 g (3) 3.50 g
(2) 0.350 g (4) 35.0 g

34. What is the percent by mass of solute in a saturated solution of $KClO_3$ at a temperature of 60°C?

(1) 0.218% (3) 21.8%
(2) 0.28% (4) 28%

35. A 2400.-gram sample of an aqueous solution contains 0.012 gram of NH_3. What is the concentration of NH_3 in the solution, expressed in parts per million?

(1) 5.0 ppm (3) 20. ppm
(2) 15 ppm (4) 50. ppm

36. Barium sulfate is listed as insoluble on solubility tables. It does, however, dissolve to a small extent. Its solubility at 50°C is 0.00034 g in 100. g of water. Express this solubility in parts per million.

37. The table below gives the solubility of $Ca(OH)_2(s)$ in 100. g of H_2O at various temperatures. What is unusual about the behavior of the solubility of this salt?

Temperature (°C)	Solubility (g $Ca(OH)_2$/100. g H_2O)
0	0.189
20	0.173
40	0.141
60	0.121
80	0.094

Colligative Properties

The freezing and boiling points of water change when nonvolatile solutes are added. When any salt is added to water, the freezing point of the water decreases. This helps explain why salt is applied to roads and sidewalks when they are covered with snow and ice. The added salt lowers the freezing point and helps to melt the snow or ice. The amount of the lowering of the freezing point is not dependent on the nature of the added particle but only on the total number of dissolved particles. One mole of any particles will have the same effect on the freezing point. One mole of particles lowers the freezing point of 1000 g of water by 1.86°C.

Molecular Versus Ionic

When one mole of sugar, a molecular substance, is dissolved in water, one mole of particles is produced in solution.

$$\begin{array}{ccc} C_{12}H_{22}O_{11}(s) & \rightarrow & C_{12}H_{22}O_{11}(aq) \\ 1 \text{ mol} & & 1 \text{ mol} \end{array}$$

When one mole of an ionic substance is dissolved in water, the results are different. The ionic substance separates into individual ions.

$$NaCl(s) \rightarrow Na^{+}(aq) + Cl^{-}(aq)$$
$$1 \text{ mol } NaCl \rightarrow 1 \text{ mol } Na^{+} + 1 \text{ mol } Cl^{-}$$

Thus, one mole of sodium chloride produces two moles of particles and will depress the freezing point of water twice as much as the mole of sugar. The greater the number of ions, the greater the effect on the freezing point. $CaCl_2$ contains three ions, and one mole of this salt will depress the freezing point three times as much as a mole of sugar.

The situation is similar with the boiling point. One mole of particles will elevate the boiling point of 1000. g of water by 0.52°C. One mole of dissolved sugar will elevate the boiling point of 1000. g of water by 0.52°C. One mole of dissolved sodium chloride contains two moles of ions, and will raise the boiling point of 1000.g of water by 1.04°C.

Review Questions Set 7.4

38. Why is salt (NaCl) put on icy roads and sidewalks in the winter?

(1) It is ionic and lowers the freezing point of water.
(2) It is ionic and raises the freezing point of water.
(3) It is covalent and lowers the freezing point of water.
(4) It is covalent and raises the freezing point of water.

39. What occurs as a salt dissolves in water?

(1) The number of ions in the solution decreases, and the freezing point decreases.
(2) The number of ions in the solution decreases, and the freezing point increases.
(3) The number of ions in the solution increases, and the freezing point decreases.
(4) The number of ions in the solution increases, and the freezing point increases.

40. Assume equal aqueous concentrations of each of the following substances. Which has the lowest freezing point?

(1) $C_2H_{12}O_6$
(2) CH_3OH
(3) $C_{12}H_{22}O_{11}$
(4) NaOH

41. What occurs when sugar is added to water?

(1) The freezing point of the water will decrease, and the boiling point will decrease.
(2) The freezing point of the water will decrease, and the boiling point will increase.
(3) The freezing point of the water will increase, and the boiling point will decrease.
(4) The freezing point of the water will increase, and the boiling point will increase.

42. Which solution has the highest boiling point?

(1) 1.0 M KNO_3
(2) 2.0 M KNO_3
(3) 1.0 M $Ca(NO_3)_2$
(4) 2.0 M $Ca(NO_3)_2$

43. Which property of a distilled water solution will not be affected by adding 50 mL of CH_3OH to 100. mL of the water solution at 25°C?

(1) conductivity
(2) mass
(3) freezing point
(4) boiling point

Memory Jogger

Hydrogen bonds are formed when a hydrogen atom in one molecule is attracted to an oxygen, nitrogen, or fluorine atom in another molecule. Because hydrogen bonding is stronger than dipole-dipole attraction, substances with hydrogen bonding have abnormally high boiling points.

Vapor Pressure

The molecules in a liquid are held together by rather weak forces. Polar molecules called dipoles are held in the liquid phase by dipole-dipole forces. In molecules containing hydrogen and one atom of oxygen, nitrogen, or fluorine, the force of attraction holding them in the liquid phase are hydrogen bonds.

In any sample of a liquid, some of the particles at the surface have sufficient energy to escape from their neighboring molecules and enter the gas phase. When a substance that is normally a solid or a liquid at room temperature enters the gas phase it is called a **vapor.** Thus, you will often hear about water vapor or gasoline vapor, as water and gasoline are normally liquids at room temperature.

As the temperature of a liquid increases, the particles have more energy, and more particles escape from the surface. These vapor particles are gaseous particles and exert pressure in the gaseous phase. The pressure that a vapor exerts is called **vapor pressure.** Table H of *Reference Tables for Physical Setting/Chemistry* is a graph showing the vapor pressure of four substances measured in pressure units of kilopascals (kPa).

Of the substances shown on the graph, propanone exerts the most pressure, about 93 kPa at a temperature of 50°C. It can be inferred that propanone has the weakest intermolecular forces holding it in the liquid phase, while ethanoic acid has the greatest, exerting only about 8 kPa of pressure.

Boiling Point

As the temperature of a liquid rises, vapor pressure increases. Finally the vapor pressure becomes equal to atmospheric pressure. At this point the gas may vaporize, not only on the surface but at any point in the container. A bubble of vapor below the surface has enough pressure that it does not collapse from the atmospheric pressure pushing against it. When a bubble can occur at any point in the liquid, the process is called boiling. The normal **boiling point** of a liquid is the temperature at which the vapor pressure of the liquid is 101.3 kPa, standard atmospheric pressure. Equivalent pressures are 1 atm, 760 mm Hg, and 760 torr. The heat required to change 1 mol of a substance from a liquid at its boiling point to 1 mol of a vapor is termed the heat of vaporization.

The normal boiling point of water is 100.°C. At this temperature, the vapor pressure of water is 101.3 kPa. The line representing 101.3 kPa on Table H shows the normal boiling point of ethanol to be 78°C. When the pressure is less than 101.3 kPa, the boiling point will be less than the normal value. Water will boil at about 70°C when the pressure is about 30 kPa. If the pressure is greater than normal, liquids will boil at temperatures above their normal boiling points. When atmospheric pressure is about 145 kPa, water boils at 110.°C.

Review Questions — Set 7.5

Use the tables in *Reference Tables for Physical Setting/Chemistry* as needed in answering the following questions.

44. What is the vapor pressure of water at 105°C?

(1) 100 kPa
(2) 101.3 kPa
(3) 110 kPa
(4) 120 kPa

45. As the pressure on a liquid is changed from 100. kPa to 120. kPa, the temperature at which the liquid will boil

(1) decreases
(2) increases
(3) remains the same

46. In a closed system at 40°C, a liquid has a vapor pressure of 50 kPa. The liquid's normal boiling point could be

(1) 10°C
(2) 30°C
(3) 40°C
(4) 60°C

47. If the pressure on the surface of water in the liquid state is 30 kPa, the water will boil at

(1) 0°C
(2) 30°C
(3) 70°C
(4) 100°C

48. A sample of ethanoic acid at 100°C has a vapor pressure of

(1) 53 kPa
(2) 100 kPa
(3) 101.3 kPa
(4) 125 kPa

49. Which substance has the greatest intermolecular forces of attraction between molecules?

(1) propanone
(2) ethanol
(3) water
(4) ethanoic acid

50. The vapor pressure of ethanol at its normal boiling point would be

(1) 80 kPa
(2) 101.3 kPa
(3) 273 kPa
(4) 373 kPa

Practice Questions
for the New York Regents Exam

Directions
Review the Test-Taking Strategies section of this book. Then answer the following questions. Read each question carefully and answer with a correct choice or response.

Part A

1 When a teaspoon of sugar is added to water in a beaker, the sugar dissolves. The resulting mixture is
(1) a compound
(2) a homogeneous solution
(3) a heterogeneous solution
(4) an emulsion

2 A small quantity of a salt is stirred into a liter of water until it dissolves. In the resulting mixture, the water is
(1) the solvent
(2) the solute
(3) dispersed material
(4) a precipitate

3 A solution
(1) will separate on standing
(2) may have color
(3) can be cloudy
(4) can be heterogeneous

4 A nonpolar solvent would most easily dissolve which of the following substances?
(1) NH_3 (3) NaCl
(2) CCl_4 (4) H_2SO_4

5 What happens when a crystal of a salt is dropped into an unsaturated solution of the same salt?
(1) Excess solute crystals form.
(2) The crystal dissolves.
(3) The crystal drops to the bottom, unchanged.
(4) The solution becomes colorless.

6 The depression of the freezing point is dependent on
(1) the nature of the solute
(2) the formula mass of the solute
(3) the concentration of dissolved particles
(4) hydrogen bonding

7 Which compound becomes *less* soluble in water as the temperature of the solution is increased?
(1) HCl (3) NaCl
(2) KCl (4) NH_4Cl

8 What happens when a crystal of solute is dropped into a supersaturated solution of the salt?
(1) The crystal dissolves.
(2) Excess solute crystals form.
(3) The crystal drops to the bottom, unchanged.
(4) The solution begins to boil.

9 Under which conditions are gases most soluble in water?
(1) high temperature and high pressure
(2) high temperature and low pressure
(3) low temperature and high pressure
(4) low temperature and low pressure

10 As the temperature of a liquid decreases, the amount of a gas that can be dissolved
(1) decreases
(2) increases
(3) remains the same

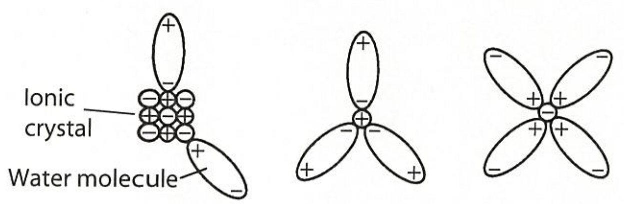
11 As the temperature of liquid water decreases, its vapor pressure
(1) decreases
(2) increases
(3) remains the same
(4) disappears

12 At standard pressure, how do the boiling point and freezing point of NaCl(*aq*) compare to that of $H_2O(\ell)$?
(1) The boiling point of NaCl(*aq*) is higher, and the freezing point of NaCl(*aq*) is lower.
(2) The boiling point of NaCl(*aq*) is lower, and the freezing point of NaCl(*aq*) is higher.
(3) Both the boiling point and freezing point of NaCl(*aq*) are lower.
(4) Both the boiling point and freezing point of NaCl(*aq*) are lower.

Part B–1

13 Which substance increases in solubility as the temperature decreases?
(1) $KClO_3$ (2) NH_3 (3) KNO_3 (4) NaCl

14 The diagrams below represent an ionic crystal being dissolved in water.



According to the diagrams, the dissolving process takes place by
(1) hydrogen bond formation
(2) metallic bonding
(3) dipole-dipole attractions
(4) molecule-ion attractions

15 If solutions of barium nitrate and sodium sulfate are mixed and then poured into a filter, the solid remaining on the filter will be
(1) barium nitrate (3) barium sulfate
(2) sodium nitrate (4) sodium sulfate

16 A student tested the solubility of a salt at different temperatures and then used the *Reference Tables for Physical Setting/Chemistry* to identify the salt. The student's data appears below.

Temperature (°C)	Solubility Data (g of salt per 10 g of water)
30	1.2
50	2.2
62	3.0
76	4.0

What is the identity of the salt?
(1) potassium nitrate (3) potassium chlorate
(2) sodium chloride (4) ammonium chloride

17 A student drops a crystal of NaCl into a beaker of NaCl(aq) and the crystal dissolves. The original solution must have been
(1) supersaturated
(2) saturated
(3) unsaturated
(4) heterogeneous

18 If 100. g of water at 80°C contains 45 g of KCl and 45 g of $NaNO_3$, the solution is
(1) saturated with respect to both KCl and $NaNO_3$
(2) saturated with respect to KCl and unsaturated with respect to $NaNO_3$
(3) unsaturated with respect to both KCl and $NaNO_3$
(4) supersaturated with respect to both KCl and $NaNO_3$

19 What happens when KI(s) is dissolved in water?
(1) I^- ions are attracted to the oxygen atoms of the water.
(2) K^+ ions are attracted to the oxygen atoms of the water.
(3) K^+ ions are attracted to the hydrogen atoms of the water.
(4) No attractions are involved; the crystal just falls apart.

20 Which compound is soluble in water?
(1) PbS (3) Na_2S
(2) BaS (4) Fe_2S_3

21 A saturated solution of potassium chloride at 10°C is heated to 30°C. As the solution is heated in a closed container, the total mass of the solution
(1) decreases (3) remains the same
(2) increases

22 Water boils at 90°C when the pressure exerted on the liquid equals
(1) 65 kPa (3) 101.3 kPa
(2) 90 kPa (4) 120 kPa

23 A solution consists of 0.50 mole of $CaCl_2$ dissolved in 100. grams of H_2O at 25°C. Compared to the boiling point and freezing point of 100. grams of pure H_2O at standard pressure, the solution at standard pressure has
(1) a higher boiling point and a lower freezing point
(2) a higher boiling point and a higher freezing point
(3) a lower boiling point and a lower freezing point
(4) a lower boiling point and a higher freezing point

24 Which sample of ethanol will have the highest vapor pressure?
(1) 10. mL at 62°C
(2) 20.0 mL at 52°C
(3) 30. mL at 42°C
(4) 40 mL at 32°C

25 The vapor pressure of water is 50 kPa. The temperature of this sample of water is
(1) 37°C (2) 62°C (3) 82°C (4) 92°C

Parts B–2 and C

Base your answers to questions 26 through 28 on the information below.

A solution of $KClO_3$ is prepared using 75 grams of the solute in enough water to make 0.250 liters of solution. The gram-formula mass of $KClO_3$ is 122 grams per mole.

26 Determine the molarity (concentration in mol/L) of the solution.

27 Determine the percent by mass of solute in the solution.

28 Assuming the solution is saturated, determine the temperature of the solution.

Base your answers to questions 29 through 31 on the information below.

Four flasks are prepared, each containing 100 mL of aqueous solution as described in the data table.

Flask	Solution	Concentration	Solute Type
A	$KCl(aq)$	1.0 M	ionic
B	$CH_3OH(aq)$	1.0 M	molecular
C	$Ba(OH)_2(aq)$	1.0 M	ionic
D	$CH_3OCH_3(aq)$	1.0 M	molecular

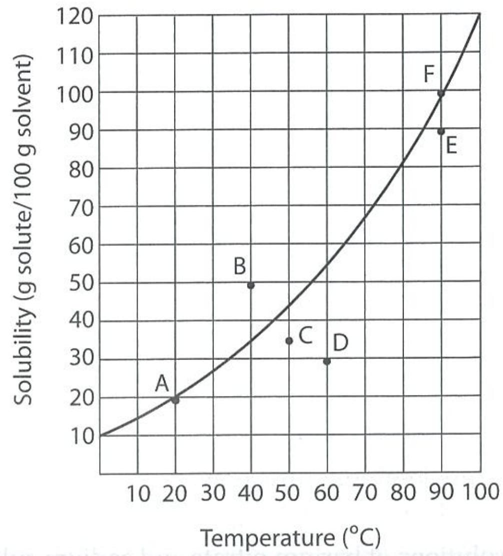
29 Which solution has the lowest freezing point?

30 For aqueous solutions, the boiling point is elevated by 0.52°C for each mole of particles (ions or molecules) present in 1 liter of the solution. Determine the boiling point of solution C.

31 Solution A is an unsaturated solution with 74.54 grams (1 mol) of solute dissolved in 1 liter of solution. According to Table G, at a temperature of 30°C, how much additional solute would be required to saturate the solution at that temperature?

Base your answers to questions 32 through 35 on the information below.

The following graph displays the Solubility Curve for substance X. The points labeled **A** through **F** represent various solutions of the substance at different concentrations and temperatures.



32 State a letter that indicates a position on this graph where solution X is saturated.

33 Describe, in terms of solute and solvent interaction, what could happen if additional solute crystals were added to the solution indicated at Point B.

34 What could be done to the solution at Point D to make the solution saturated?

35 If the gram-formula mass of substance X is 180 g/mol, determine the molarity of the solution at Point E.

Base your answers to questions 36 through 38 on the following information.

A scientist makes a solution that contains 44.0 grams of hydrogen chloride gas, $HCl(g)$, in 200. grams of water, $H_2O(\ell)$, at 20.°C. This process is represented by the balanced equation below.

$$HCl(g) \xrightarrow{H_2O} H^+(aq) + Cl^-(aq)$$

36 Based on Reference Table G, identify, in terms of saturation, the type of solution made by the scientist.

37 Explain, in terms of distribution of particles, why the solution is a homogeneous mixture.

38 Knowing that the gram-formula mass of HCl is 36 g/mol, determine the molarity of this solution.