Chapter 16
"*Solutions*"

Honors Chemistry

Section 16.1
Properties of Solutions

- OBJECTIVES:
  - Identify the factors that determine the ***rate*** at which a solute *dissolves*.

Section 16.1
Properties of Solutions

- OBJECTIVES:
  - Identify the *units* usually used to express the *solubility* of a solute.

Section 16.1
Properties of Solutions

- OBJECTIVES:
  - Identify the factors that determine the mass of solute that *will dissolve* in a given mass of solvent.

Solution formation

- The “nature” (polarity, or composition) of the solute and the solvent will determine…

1. Whether a substance *will dissolve*
2. How *much* will dissolve

- Factors determining **rate** of solution...

3. stirring (agitation)
4. surface area the dissolving particles
5. temperature

Making solutions

- In order to dissolve, the solvent molecules must come *in contact* with the solute.

1. Stirring (agitation) moves fresh solvent into contact with the solute.
2. Smaller pieces increase the amount of surface area of the solute.

- think of how fast a breath mint dissolves when you chew it

Temperature and Solutions

3. Higher temperature makes the molecules of the solvent *move faster* and contact the solute harder and more often.
   - Speeds up dissolving.

- *Higher Temperature* ***ALSO*** *Usually* increases the amount that will dissolve (an exception is gases, more on that later).

Figure 16.4 Interpreting Graphs - Page 474

a. The solubility of the KNO_3 increases as the temperature increases.

b. $Yb_2(SO_4)_3$ shows a decrease in solubility as the temperature increases, and NaCl shows the least change in solubility as temperature changes.

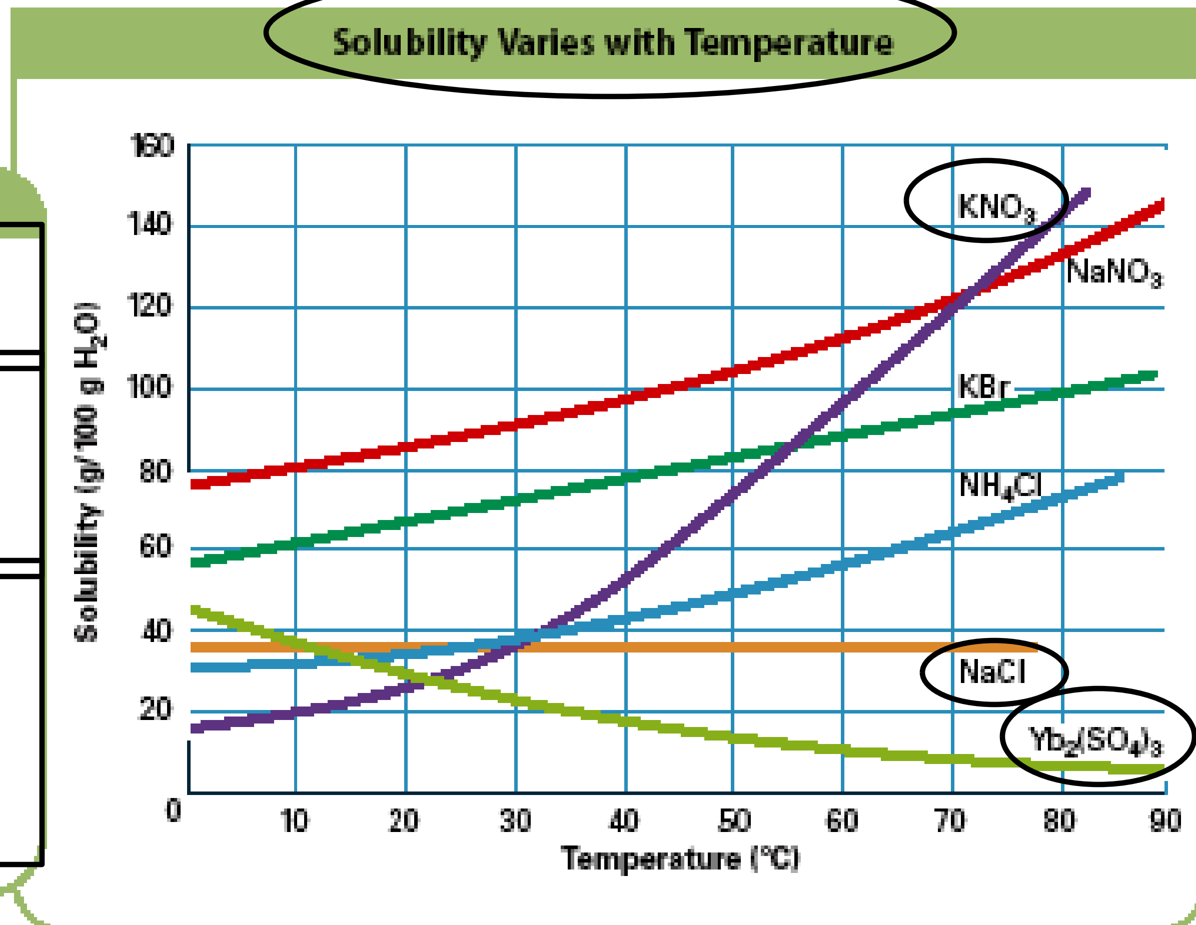
c. Only a negligible amount of NaCl would go into solution, if any.

INTERPRETING GRAPHS

a. **Describe** What happens to the solubility of KNO_3 as the temperature increases?

b. **Identify** Which substance shows a decrease in solubility as temperature increases? Which substance exhibits the least change in solubility?

c. **Apply Concepts** Suppose you added some solid sodium chloride (NaCl) to a saturated solution of sodium chloride at 20°C and warmed the mixture to 40°C. What would happen to the added sodium chloride?



Solids tend to dissolve best when:

- They are heated
- They are stirred
- Crushed into smaller particles

Gases tend to dissolve best when:

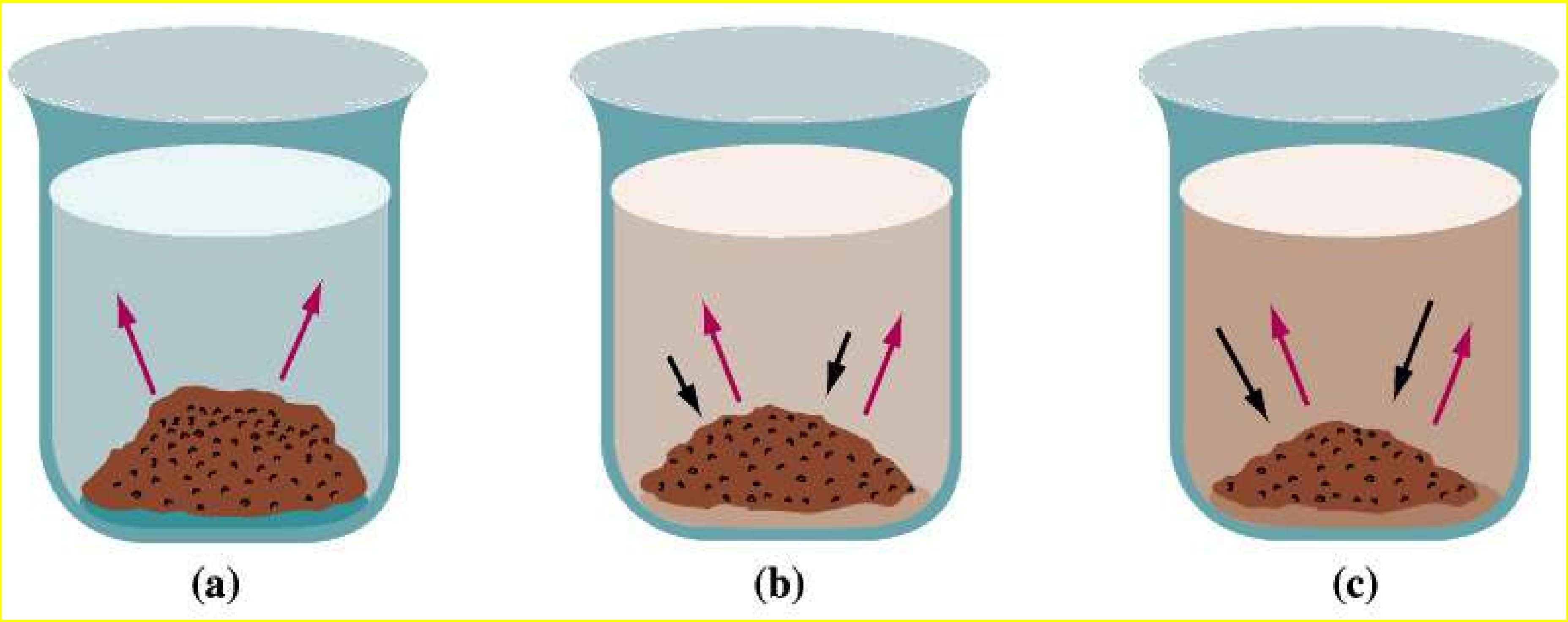
- The solution is cold
- The pressure is high

How Much?

- **Solubility**- is the maximum amount of substance that will dissolve at a specific temperature. The units for solubility are: grams of solute/100 grams solvent

1) **Saturated solution**- Contains the *maximum* amount of solute dissolved. NaCl = 36.0 g/100 mL water
2) **Unsaturated solution**- Can still dissolve more solute (for example 28.0 grams of NaCl/100 mL)
3) **Supersaturated**- solution that is holding (or dissolving) more than it theoretically can; a “seed crystal” will make it come out; Fig. 16.6, page 475

Saturation and Equilibrium

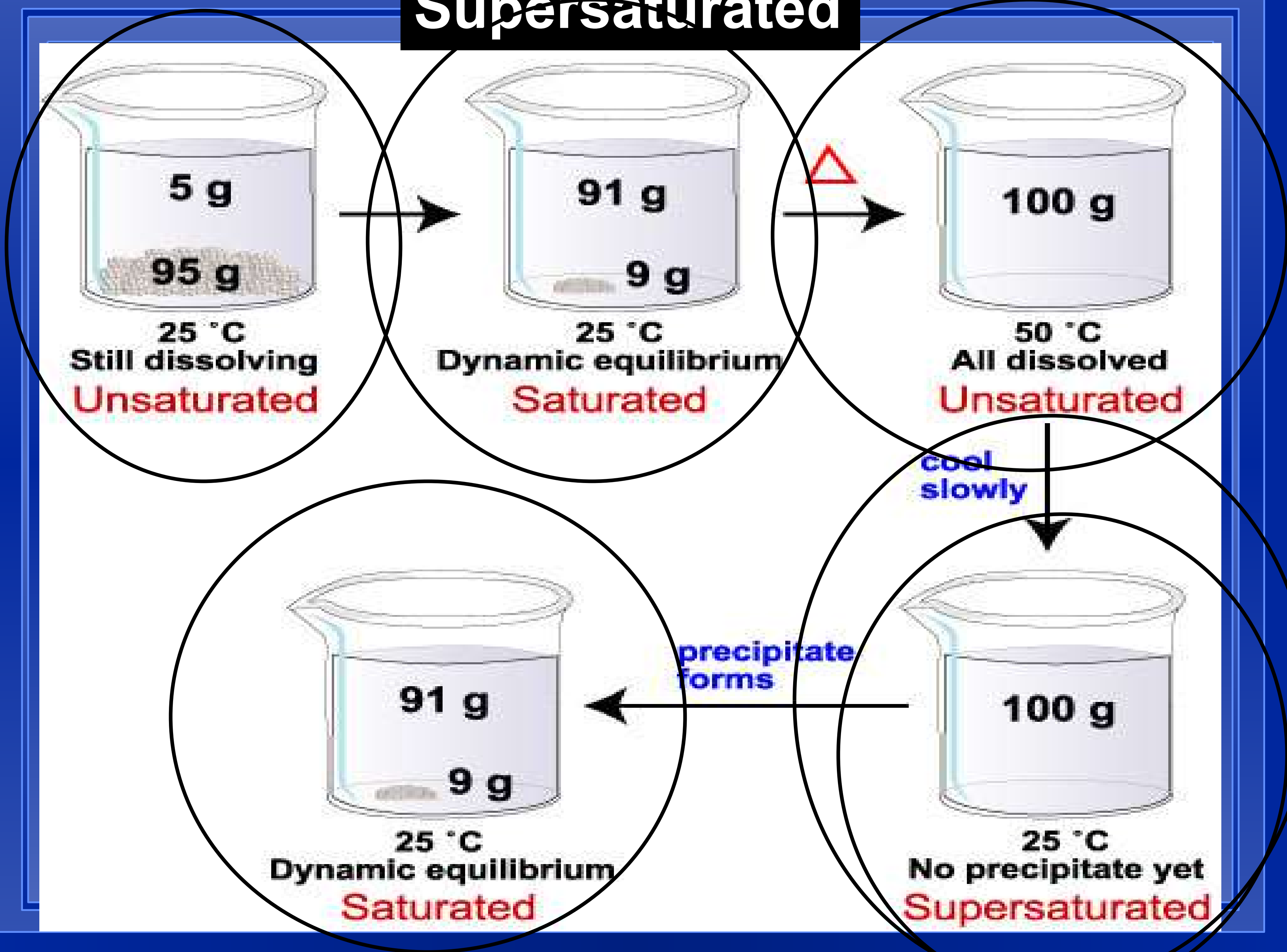


Solute is dissolvin g

More solute is dissolving, but some is crystallizing

Saturation equilibriu m establishe

Supersaturated



Supersaturated Example

- Ever heard of "seeding" the clouds to make them produce rain?
- Clouds - mass of air supersaturated with water vapor
- Silver Iodide (AgI) crystals are dusted into the cloud as a "seed"
- The AgI attracts the water, forming droplets that attract others

Liquids

- Miscible means that two liquids can dissolve in each other
  - water and antifreeze
  - water and ethanol
- Partially miscible- slightly
  - water and ether
- Immiscible means they can't
  - oil and vinegar

Solubility?

- For solids in liquids, as the temperature goes up-the solubility usually goes up (Fig. 16.4, p.474)
- For gases in a liquid, the effect is the opposite of solids in liquids
  - As the temperature goes up, gas solubility goes down
  - Think of boiling water bubbling?
  - Thermal pollution may result from industry using water for cooling

Gases in liquids...

- Henry's Law - says the solubility of a gas in a liquid is *directly proportional* to the pressure of the gas *above the liquid*
  - think of a bottle of soda pop, removing the lid releases pressure
- Equation:

$$\frac{S_1}{P_1} = \frac{S_2}{P_2}$$

Sample 16.1, page 477

Section 16.2
Concentration of Solutions

- OBJECTIVES:
  - Solve problems involving the molarity of a solution.

Section 16.2
Concentration of Solutions

- OBJECTIVES:
  - Describe the effect of dilution on the total moles of solute in solution.

Section 16.2
Concentration of Solutions

- OBJECTIVES:
  - Define percent by volume and percent by mass solutions.

Concentration is...

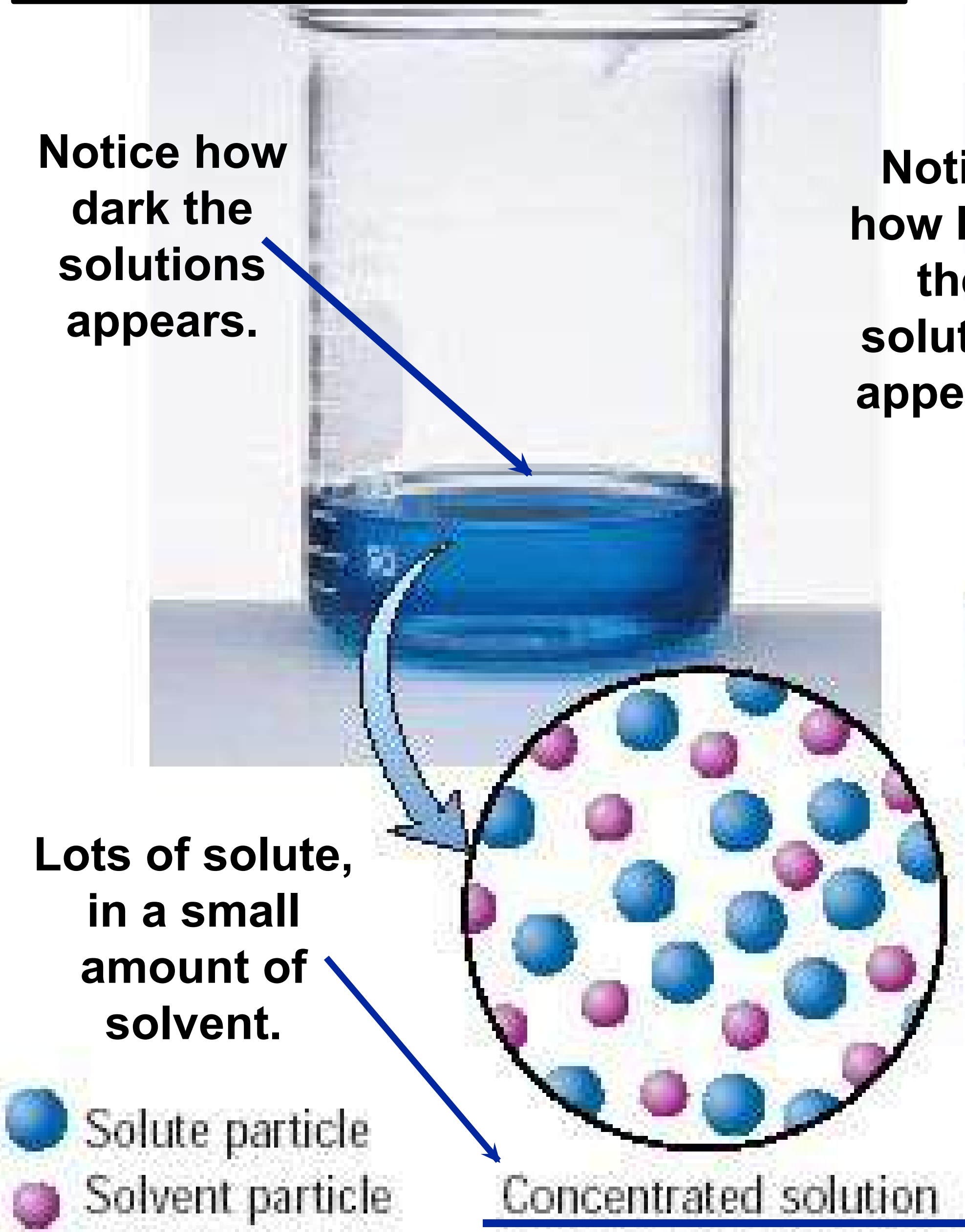
- a measure of the amount of solute dissolved in a given quantity of solvent
- A ***concentrated*** solution has a *large* amount of solute
- A ***dilute*** solution has a *small* amount of solute
  - These are *qualitative* descriptions
- But, there are ways to express solution concentration *quantitatively (NUMBERS!)*

CONCENTRATED

DILUTE

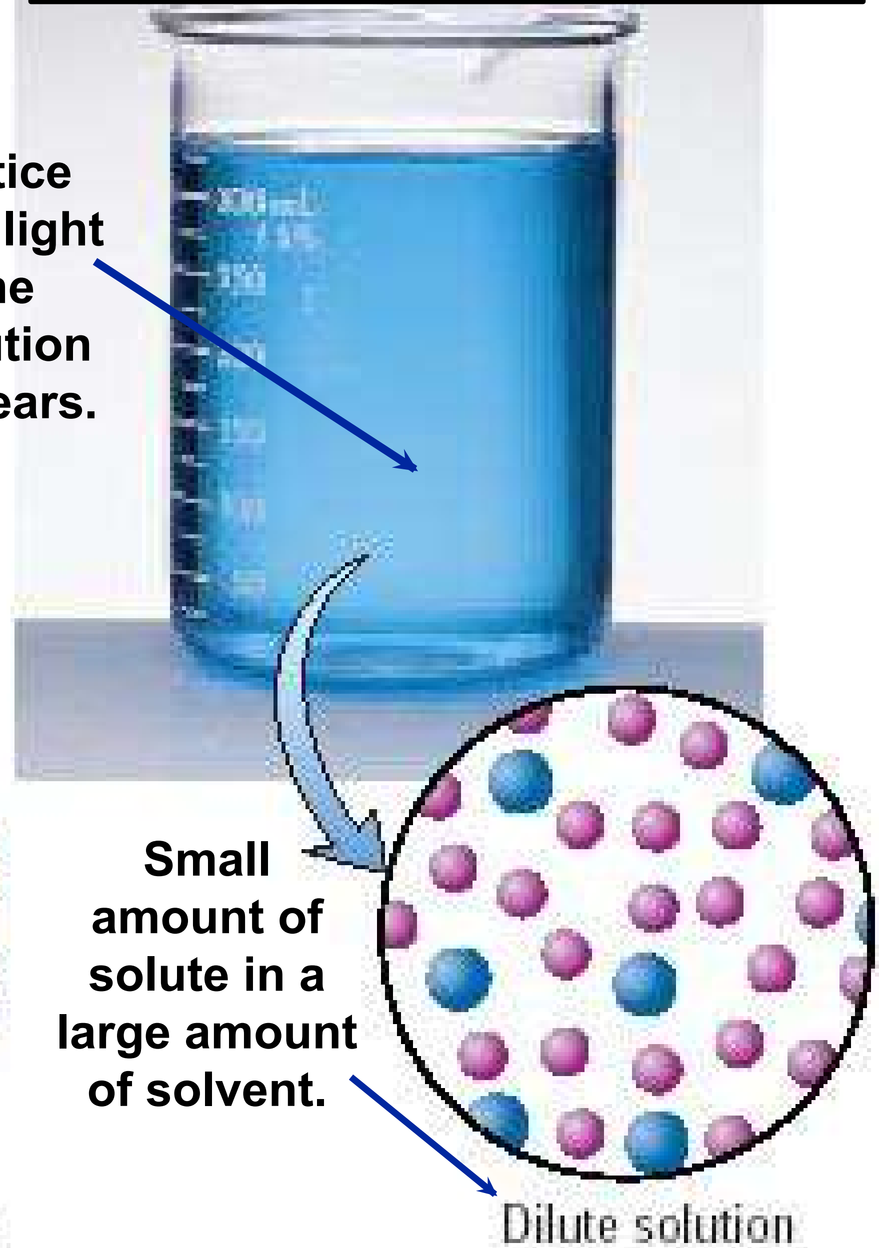
Notice how dark the solutions appears.

Lots of solute, in a small amount of solvent.



Notice how light the solution appears.

Small amount of solute in a large amount of solvent.



Molarity: a unit of concentration

- Molarity = $\frac{\text{moles of solute}}{\text{liters of solution}}$
  - Abbreviated with a capital M, such as 6.0 M
- This is the most widely used concentration unit used in chemistry.

SAMPLE PROBLEM 16.2 - Page 481

Calculating the Molarity of a Solution

Intravenous (IV) saline solutions are often administered to patients in the hospital. One saline solution contains 0.90 g NaCl in exactly 100 mL of solution. What is the molarity of the solution?

1 Analyze ***List the knowns and the unknown.***

Knowns

- solution concentration = 0.90 g NaCl/100 mL
- molar mass NaCl = 58.5 g/mol

Unknown

- solution concentration = ?M

Convert the concentration from g/100 mL to mol/L. The sequence is g/100 mL → mol/100 mL → mol/L.

2 Calculate ***Solve for the unknown.***

Use the molar mass to convert g NaCl/100 mL to mol NaCl/100 mL. Then use the conversion factor between milliliters and liters to convert to mol/L, which is molarity.

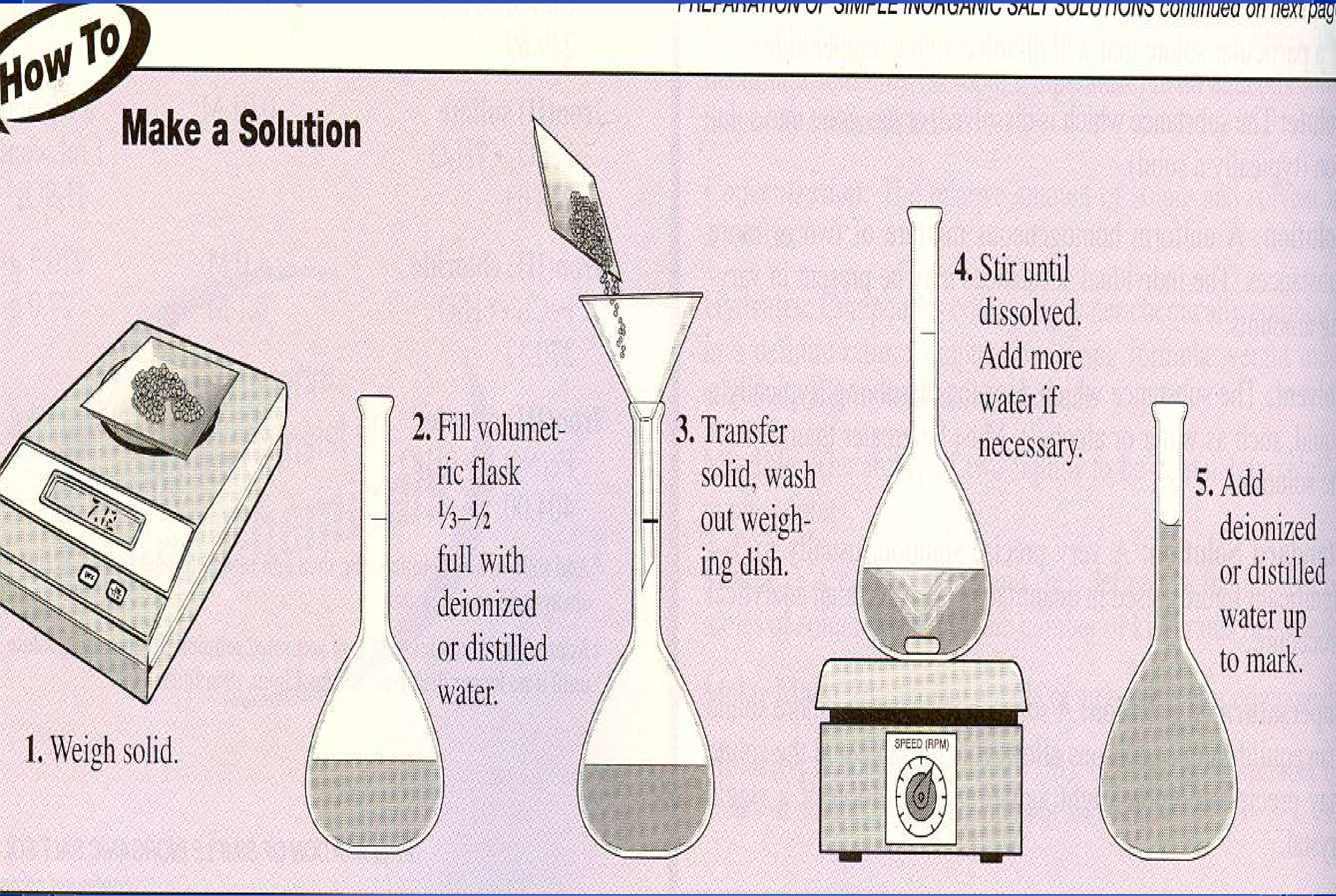
$$\text{solution concentration} = \frac{0.90\ \cancel{\text{g NaCl}}}{100\ \cancel{\text{mL}}} \times \frac{1\ \text{mol NaCl}}{58.5\ \cancel{\text{g NaCl}}} \times \frac{1000\ \cancel{\text{mL}}}{1\ \text{L}}$$

$$= 0.15\ \text{mol/L} = 0.15M$$

Making solutions

1) Pour in a *small amount* of the solvent, maybe about one-half
2) Then add the pre-massed solute (and mix by swirling to dissolve it)
3) Carefully fill to final volume.
   - Fig. 16.8, page 481, and shown on next slide.

- Can also solve: moles = M x L
- Sample Problem 16.3, page 482


PREPARATION OF SIMPLE INORGANIC SALT SOLUTIONS continued on next page
How To
Make a Solution
1. Weigh solid.
2. Fill volumetric flask ⅓–½ full with deionized or distilled water.
3. Transfer solid, wash out weighing dish.
4. Stir until dissolved. Add more water if necessary.
5. Add deionized or distilled water up to mark.
SPEED (RPM)

Dilution

- Adding water to a solution will reduce the number of moles of solute *per unit volume*
  - but the overall number of moles remains the same!
- Think of taking an aspirin with a small glass of water vs. a large glass of water
  - You still have *one aspirin* in your body, regardless of the amount of water you drank, but a larger amount of water makes it more diluted.

Dilution

- The number of moles of solute in solution doesn't change if you add more solvent!
- The # moles before = the # moles after
- Formula for dilution: $M_1 \times V_1 = M_2 \times V_2$
- M_1 and V_1 are the starting concentration and volume; M_2 and V_2 are the final concentration and volume.
- *Stock solutions* are pre-made solutions to known Molarity. Sample 16.4, p.484

Percent solutions can be expressed by a) volume or b) mass

- Percent means parts per 100, so
- Percent by volume: = Volume of solute ~~x 100% Volume~~ of solution
- indicated %(v/v)
- Sample Problem 16.5, page 485

Percent solutions

- Percent by mass: = Mass of solute(g) x 100% Volume of solution (mL)
- Indicated %(m/v)
- More commonly used
- 4.8 g of NaCl are dissolved in 82 mL of solution. What is the percent of the solution?
- How many grams of salt are there in 52 mL of a 6.3 % solution?

Percent solutions

- Another way to do mass percentage is as mass/mass:

- Percent by mass: = Mass of solute(g) x 100% Mass of solution (g)

- Indicated %(m/m)

Section 16.3 Colligative Properties of Solutions

- OBJECTIVES:
  - Identify three *colligative properties* of solutions.

Section 16.3
Colligative Properties of Solutions

- OBJECTIVES:
  - Explain why the vapor pressure, freezing point, and boiling point of a solution *differ* from those properties of the pure solvent.

Colligative Properties

-These depend only on the ***number*** of dissolved particles

-Not on what ***kind*** of particle

-Three important colligative properties of solutions are:

1) Vapor pressure *lowering*
2) Boiling point *elevation*
3) Freezing point *lowered*

Some particles in solution will IONIZE (or split), while others may not.

Figure 16.14 Particle Concentrations in Solutions

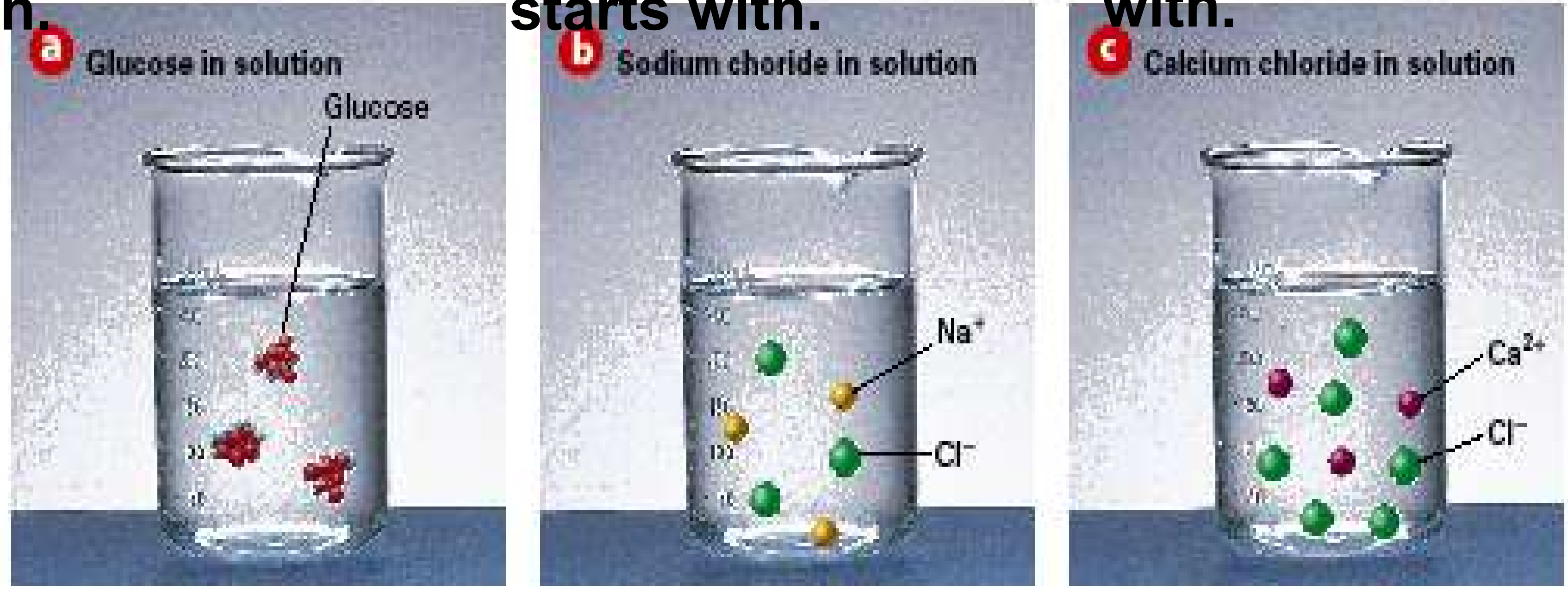
- Page 488

Colligative Properties

Glucose will only have one particle in solution for each one particle it starts with.

NaCl will have two particles in solution for each one particle it starts with.

$CaCl_2$ will have three particles in solution for each one particle it starts with.



Vapor Pressure is LOWERED

1) Surface area is reduced, thus less evaporation, which is a surface property
2) The bonds between molecules keep molecules from escaping. So, in a solution, some of the solvent is busy keeping the solute dissolved.

- This **lowers** the vapor pressure
- Electrolytes *form ions* when they are dissolved, making more pieces.
- $NaCl \rightarrow Na^+ + Cl^-$ (this = 2 pieces)
- More pieces = a bigger effect

Boiling Point is ELEVATED

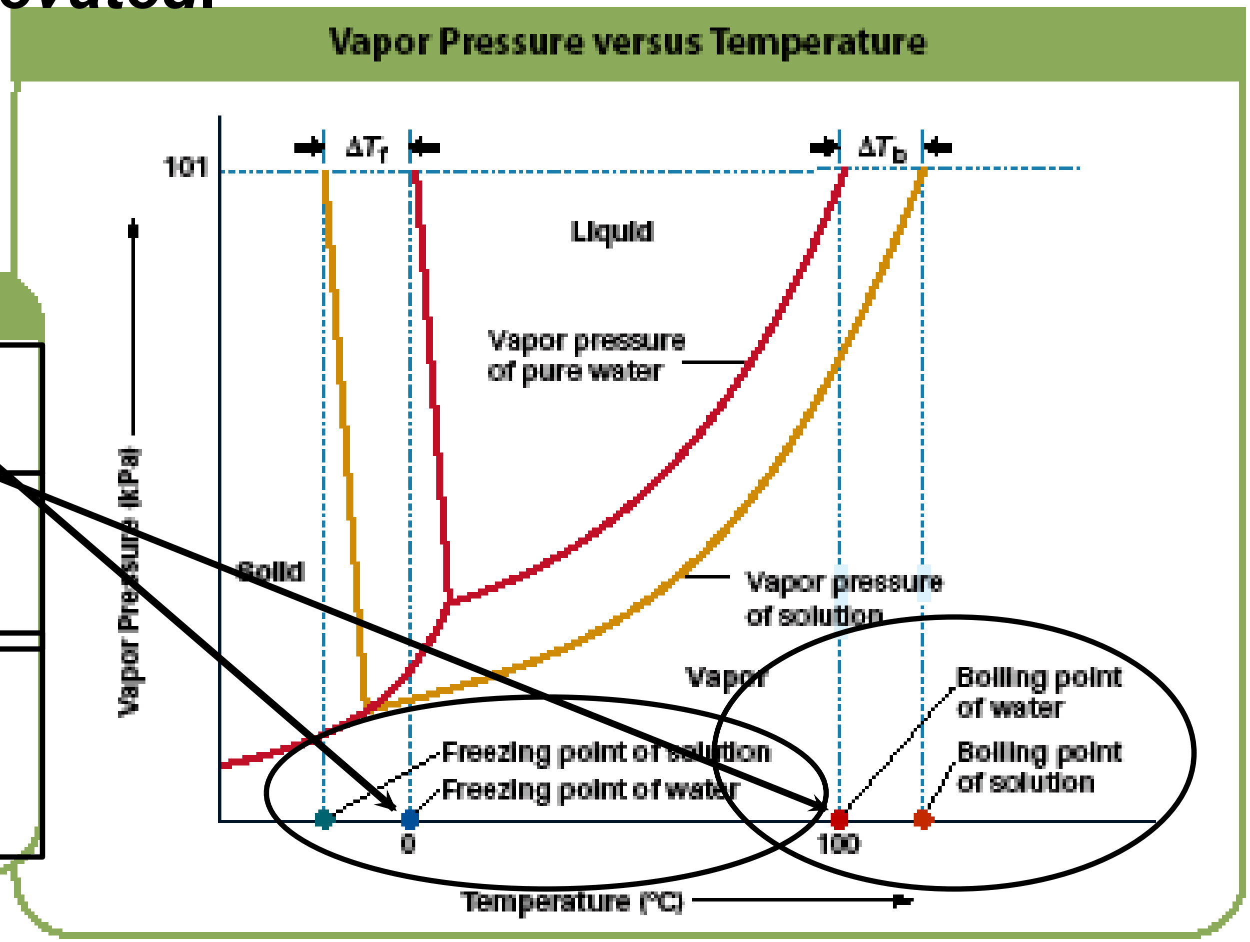
- The vapor pressure determines the boiling point. (Boiling is defined as when the vapor pressure of liquid = vapor pressure of the atmosphere).
- Lower vapor pressure means you need a *higher temperature* to get it to equal atmospheric pressure
- Salt water boils above 100ºC
- The number of dissolved particles determines how much, as well as the solvent itself.

Freezing Point is LOWERED

- Solids form when molecules make an orderly pattern called “crystals”
- The solute molecules break up the orderly pattern.
  - Makes the freezing point lower.
  - Salt water freezes below 0ºC
  - Home-made ice cream with rock salt?
- How much lower depends on the amount of solute dissolved.

Figure 16.20 Interpreting Graphs - Page 494

The addition of a solute would allow a <u>LONGER</u> temperature range, since freezing point is *lowered* and boiling point is *elevated*.



INTERPRETING GRAPHS

a. Identify What is the freezing point of water? What is the boiling point?

b. Describe How do the freezing and boiling points of the solution compare to those of pure water?

c. Apply Concepts Does adding a solute to water allow it to remain as a liquid over a longer or shorter temperature range? Explain.

Section 16.4 Calculations Involving Colligative Properties

- OBJECTIVES:
  - Solve problems related to the molality and mole fraction of a solution.

Section 16.4
Calculations Involving Colligative Properties

- OBJECTIVES:
  - Describe how freezing point depression and boiling point elevation are related to molality.

Molality (abbreviated m)

- a new unit for concentration
- *m* = Moles of solute / kilogram of solvent
- *m* = Moles of solute / 1000 g of solvent
- Sample Problem 16.6, p. 492

Mole fraction

- This is another way to express concentration
- It is the ratio of moles of solute to total number of moles of solute plus solvent (Fig. 18-19, p.522)

$$X = \frac{n_a}{n_a + n_b}$$

Sample 16.7, page 493

Freezing Point Depression

- The size of the change in freezing point is also determined by molality.

$\Delta T_f = -K_f \times m \times n$

- ΔT_f is the change in freezing point
- K_f is a constant determined by the solvent (Table 16.2, page 494).
- n is the number of pieces it falls into when it dissolves.

SAMPLE PROBLEM 16.8 - Page 495

Calculating the Freezing-Point Depression of a Solution

Antifreeze protects a car from freezing. It also protects it from overheating. Calculate the freezing-point depression and the freezing point of a solution containing 100 g of ethylene glycol ($C_2H_6O_2$) antifreeze in 0.500 kg of water.

1 Analyze *List the knowns and the unknown.*

Knowns
- mass of solute = 100 g $C_2H_6O_2$
- mass of solution = 0.500 kg H_2O
- K_f for H_2O = 1.86°C/m
- $\Delta T_f = K_f \times m$

Unknown
- ΔT_f = °C
- freezing point = ?°C

Calculate the number of moles of solute and the molality. Then calculate the freezing-point depression and freezing point.

2 Calculate *Solve for the unknown.*

$$\text{moles } C_2H_6O_2 = 100\ \cancel{\text{g } C_2H_6O_2} \times \frac{1 \text{ mol}}{62.0\ \cancel{\text{g } C_2H_6O_2}} = 1.61 \text{ mol}$$

$$m = \frac{\text{mol solute}}{\text{kg solvent}} = \frac{1.61 \text{ mol}}{0.500 \text{ kg}} = 3.22m$$

$$\Delta T_f = K_f \times m = 1.86°\text{C}/m \times 3.22m = 5.99°\text{C}$$

The freezing point of the solution is 0.00°C − 5.99°C = −5.99°C.

Boiling Point Elevation

- The size of the change in boiling point is determined by the molality.

$\Delta T_b = K_b \times m \times n$

- ΔT_b is the change in the boiling point
- K_b is a constant determined by the solvent (Table 16.3, page 495).
- n is the number of pieces it falls into when it dissolves.
- Sample Problem 16.9, page 496

Key Equations

- Note the key equations on page 498 to solve problems in this chapter.

End of Chapter 16