

## Chapter 12: Solutions

Problems: 3, 5, 8, 12, 14, 16, 22, 29, 32, 41-58, 61-68, 71-74

**solution**: homogeneous mixture of a **solute** dissolved in a **solvent**

**solute**: component present in *smaller* amount

**solvent**: component present in *greater* amount

**solubility**: Maximum amount of solute dissolved in solvent at specific temp. (T)  
– In units of **g solute/100 g solvent** or **g solute/100 mL solvent**

### 12.2 Solubility and the Solution Process

#### Unsaturated, Saturated, and Supersaturated Solutions

**unsaturated**: contains *less than the maximum* amount of solute that a solvent can hold at specific temperature

**saturated**: contains *the maximum* amount of solute that a solvent can hold at specific temperature

**supersaturated**: contains *more than the maximum* amount of solute that a solvent can hold at specific temperature

How? At higher temperature, solvents can hold more solute than at lower temperature. If a given amount of solute is dissolved in a solvent at a higher temperature, then allowed to cool without being disturbed, the solute will remain in solution.

The solution is unstable, though, and the solute will fall out of solution if the solution is disturbed.

This is the principle behind hot and cold packs. Hot packs made from substances that recrystallize in an exothermic reaction; cold packs made from substances that recrystallize in an endothermic reaction.

#### Factors in Explaining Solubility

**Three types of interactions to consider for solutions:**

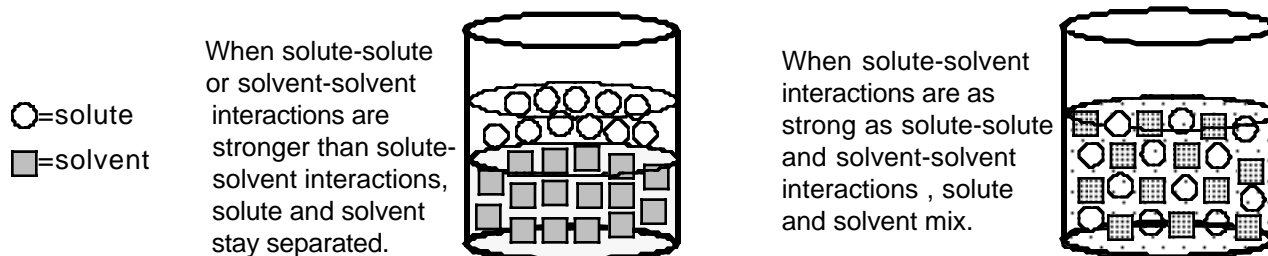
1. solvent-solvent interaction
2. solute-solute interaction
3. solvent-solute interaction

Consider the solution process taking place in three distinct steps:

1. separation of solvent molecules
2. separation of solute molecules
3. mixing of solvent and solute molecules

**So why don't all liquids mix and all solids dissolve in liquids?**

- If solvent-solute interaction can't compete with solute-solute and solvent-solvent interactions, then remain separated.



## Liquid-Liquid Solutions

**"Like dissolves like" rule**

- polar molecules will mix (be **miscible** with) other polar molecules
- nonpolar molecules will mix (be **miscible** with) other nonpolar molecules
- polar molecules will not mix (be **immiscible** with) nonpolar molecules

## Solid-Liquid Solutions

**Ionic and Molecular Compounds**

- "Like dissolves like" rule applies!

solid solute	polar solvent	nonpolar solvents
polar	soluble	insoluble
nonpolar	insoluble	soluble
ionic	Check Solubility Rules	insoluble

**Note:** You need to be able to determine whether a compound is polar or nonpolar given only its formula—i.e. get the Lewis structure, and use VSEPR to determine shape, then polarity.

**Some solids will not dissolve in any solvent**

- **network covalent solids** (eg. graphite, quartz) never dissolve in any solvent
- **metals** do not "dissolve"—they may react with but do not dissolve—in solvents

Example 1: Which of the following will be soluble in or miscible with water?  
(Circle all that apply)

$I_2$       NaCl       $Mg(OH)_2$        $Br_2(l)$        $NH_3$        $C_{diamond}$

Example 2: Which of the following will be soluble in or miscible with hexane,  $C_6H_{14}$ , a nonpolar liquid?(Circle all that apply)

$I_2$       NaCl       $Mg(OH)_2$        $Br_2(l)$        $NH_3$        $C_{diamond}$

### 12.3 Effects of Temperature and Pressure on Solubility

#### Gas solubility and Temperature:

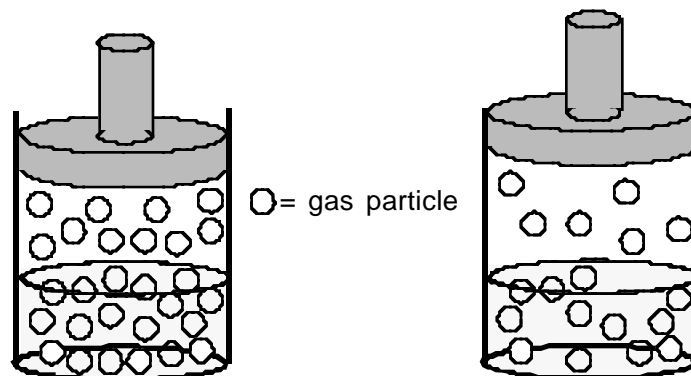
- As  $T \uparrow$ , solubility of a gas in a liquid  $\downarrow$  (in most cases)
- a glass of soda quickly goes flat if left out on a hot summer day
- why bubbles form when water heated in open pan (dissolved air escaping)

Why? At higher  $T$ , gas molecules are moving more quickly  
→ they have a higher tendency to find the surface (b/w liquid and air)  
→ they escape more quickly  
→ fewer gas molecules in the liquid!

#### Gas solubility and Pressure:

Henry's Law: Solubility of gas is proportional to partial  $P$  of gas above liquid

Why? Greater gas pressure = more gas molecules over solution  
→ more gas molecules encounter liquid surface  
→ more gas molecules go into the liquid phase!



Practical applications:

- Why sodas can't be as carbonated after being opened
- Why divers get the "bends" – air dissolved in blood and other bodily fluids bubbles out when divers go from deep water (high pressure) to the surface (low pressure)

### Solid solubility and Temperature:

- As  $T \uparrow$ , solubility of a solid in a liquid  $\uparrow$  (in most cases)
- e.g. we can dissolve more sugar in a cup of hot tea than in glass of iced tea

## 12.4 Ways of Expressing Concentration

Mass Percent Concentration (M/M%) of Solute; Parts per Million

$$\mathbf{M/M\%} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\% = \frac{\text{mass of solute}}{\text{mass of solute} + \text{mass of solvent}} \times 100\%$$

Example: A 1.215-g sample of NaCl is dissolved in 65.483 g of water.  
What is the mass percent of NaCl in the solution?

$$\mathbf{\text{parts per million (ppm) of solute}} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 10^6$$

Example: A 2.500-g sample of groundwater was found to contain 5.4 micrograms ( $\mu\text{g}$ ) of  $\text{Zn}^{2+}$ . What is the concentration of  $\text{Zn}^{2+}$  in parts per million?

**Mole Fraction (X):** has no units since ratio of two similar quantities

$$\mathbf{\text{mole fraction of component A}} = \mathbf{X_A} = \frac{\text{moles of A}}{\text{total number of moles for all components}}$$

**Molarity (or Molar Concentration)** =  $\frac{\text{moles of solute}}{\text{liters (L) of solution}}$ , in units of **molar, M**

**Molality:** number of moles of solute dissolved in 1 kg (1000 g) of solvent

$$\text{Molality (or Molal Concentration)} = \frac{\text{moles of solute}}{\text{kg of solvent}} \text{ in units of } \textit{molal} = m$$

Example: Calculate the molality of a sulfuric acid (MW=98.086 g/mol) solution containing 25.6 g of sulfuric acid in 195 g of water.

### Comparison of Concentration Units:

mole fractions: used for partial pressures of gases and for dealing with vapor pressures of solutions

molarity: preferred over molality because easier to measure volume of a solution using calibrated glassware than to weigh solvent

molality: independent of temperature, whereas molarity varies with temperature since volume varies  
– useful when experiment carried out over a range of temperatures

mass percent: independent of temperature; molar masses not needed

parts per million: for very low concentrations of solute (impurities, pollutants)

**Be able to use unit analysis to convert from one concentration unit to another!**

### Colligative Properties

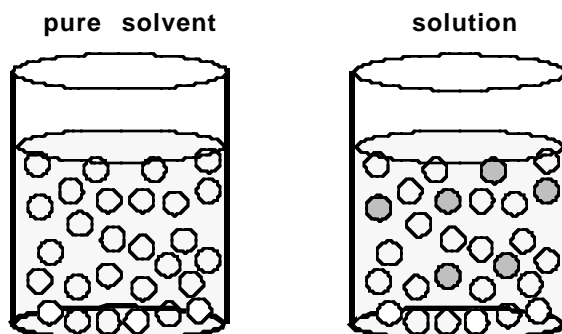
**colligative properties:** properties depending on the number of solute particles in solution and not on the nature of the solute particles

**nonelectrolytes:** exist as molecules in solution (do not dissociate into ions)

**electrolytes:** exist as ions in solution

## 12.5 Vapor Pressure Lowering (Nonelectrolytes)

- Adding a solute lowers the concentration of solvent molecules in liquid phase since solute particles block solvent molecules from going to gas phase



**Raoult's Law:**  $P_{\text{solution}} = P_{\text{solvent}} X_{\text{solvent}}$  where  $P_{\text{solution}}$  = partial P of solution  
 $P_{\text{solvent}}$  = partial P of pure solvent  
 $X_{\text{solvent}}$  = mole fraction of solvent

## 12.6 Boiling-Point Elevation and Freezing Point Depression (Nonelectrolytes)

### Boiling-Point Elevation:

- presence of solute lowers vapor pressure of a solution (since solute particles present at interface block solvent molecules from escaping)
- e.g. add salt to water to increase boiling point of water

$$T_b = T_b^\circ + \Delta T_b$$

where  $T_b$ =boiling point of solution,  $T_b^\circ$ =boiling point of pure solvent, and  $\Delta T_b$ =boiling point elevation

- $\Delta T_b$  can be calculated using  $\Delta T_b = K_b m$   
where  $m$  = molal concentration of solute,  $K_b$  = molal boiling point constant

### Freezing-Point Depression:

- the amount of impurity (or solute present) determines how much freezing point is changed
- e.g. adding salt to roads/sidewalks to prevent them from freezing in winter

$$T_f = T_f^\circ - \Delta T_f$$

where  $T_f$ =freezing point of solution,  $T_f^\circ$ =freezing point of pure solvent, and  $\Delta T_f$ =freezing point depression

- $\Delta T_f$  can be calculated using  $\Delta T_f = K_f m$   
where  $m$  = molal concentration of solute,  $K_f$  =molal freezing point constant

Ex. 1: Ethylene glycol (EG) is a commonly used automotive antifreeze. A 25.0% solution of EG is used. What is the molality of this solution? Molar mass of EG=60.052 g/mol.

Ex. 2 Calculate the boiling point and the freezing point of the solution if  $K_b=0.512\text{ }^\circ\text{C/m}$  and  $K_f=1.858\text{ }^\circ\text{C/m}$  for water.

## Osmotic Pressure (Nonelectrolytes)

**semipermeable membrane:** allows solvent molecules to pass through but blocks the passage of solute molecules

**osmosis:** net movement of solvent molecules through a semipermeable membrane from a pure solvent or a more dilute solution to a more concentrated solution  
–result of differences in vapor pressure due to differences in concentrations

Example: We often hear the phrase "learning by osmosis." If a person's head was a semipermeable membrane and we consider knowledge the solute. What would literally happen to the brain if a person was actually able to learn by osmosis?

**osmotic pressure ( $\pi$ ):** pressure required to stop osmosis

**isotonic:** when two solutions have equal osmotic pressure

**hypertonic:** the more concentrated solution given two solutions with different osmotic pressures

**hypotonic:** the less concentrated solution given two solutions with different osmotic pressures

If we make the external pressure large enough (i.e. for  $P > \pi$ ), we can make water molecules go from solution to pure water. This process is called **reverse osmosis** and is often used to obtain drinking water.

## Colligative Properties of Electrolyte Solutions

For colligative properties,

- Electrolyte solutions vary from those of nonelectrolyte solutions
  - electrolytes dissociate into ions in solutions
    - increasing the total number of particles in solution
    - freezing-point depression for 0.1 *m* NaCl is nearly double that for 0.1 *m* sucrose

**van't Hoff factor =  $i$**  = number of moles of ions per mole of electrolyte

$$i = \frac{\text{actual number of particles in solution after dissolving}}{\text{number of compounds involved before dissolving}}$$

Example: Give  $i$  for each of the following solutions:

a.  $\text{Na}_3\text{PO}_4$ :  $i =$

b.  $\text{CaCl}_2$  and  $\text{KOH}$ :  $i =$

c.  $\text{Al}(\text{NO}_3)_3$  and  $\text{H}_2\text{SO}_4$ :  $i =$



For electrolyte solutions, the van't Hoff factor can be included in the following:

$$\Delta T_b = i K_b m \quad \text{and} \quad \Delta T_f = i K_f m$$

We can also use these equations to determine the van't Hoff factor for a solution.

Ex. 1: Estimate the boiling point of a  $2.50m$  solution of  $MgCl_2$  if  $K_b=0.512 \text{ }^\circ\text{C}/m$  for water.

Ex. 2: Which of the following solutions would have the highest boiling point? (Circle one)

$1.00m$  NaBr

$1.00m$   $CaBr_2$

$1.00m$   $AlBr_3$

$1.00m$   $SrSO_4$