**Unit 10: Chapters 15 and 16**

Water, Aqueous Systems, Solutions, and Solution Formation

**Chapter 15**

1. Properties of Water
   1. Covalent bonds and polarity:
   2. Hydrogen bonding: bonding that occurs between water molecules.
   3. Low vapor pressure (h-bonding make it more difficult to change from a liquid to a gas.
   4. High boiling point
   5. High Specific heat
   6. Surfactant: surface acting agent that breaks apart hydrogen bonds

2. Phase changes of water: Ice ↔ water ↔ steam

a. Density-water is the most dense at 4 ºC: **1 gram /1 milliliter**

b. Ice: a honeycomb structure due to h-bonding

3. **Homogeneous Solutions:** no visible particles when mixed.

a. Solute: dissolved particles

b. Solvent: the dissolving medium

*Examples:* gas in solid, solid in solid, gas in gas, gas in liquid, etc.

c. Aqueous Solution: Water is the dissolving medium

*Example:* Like dissolves like

\* charged dissolves charges: polar/polar(coval. unequal sharing)

polar/ionic (metal-nonmetal bonds)

ionic/ionic

\*non charged dissolves non charged: nonpolar/nonpolar

d. Solvation: process that occurs when a solute dissolves

*Example:* NaCl (s) → Na +(aq) + Cl -(aq)

The salt breaks apart andgives two ions.

Example: MgBr2 (s) → Mg +2(aq) + 2Br -(aq)

This salt breaks apart into three ions.

e. Electrolytes

\* solutions that conduct electricity (molten or aqueous)

\* form ions in solution

\*all ionic compounds including acids and bases

\* strong electrolytes ionize completely, weak ionize partially, and non electrolytes do not ionize. The more ions formed, the better the conductivity.

f. Water of Hydration: water in a crystal and is called a hydrated crystal.

Example: CuSO4\*5H2O (s) → CuSO4(s) + 5H2O(g)

\* *Effloresce*: crystals that lose water or humidify the air.

Vapor pressure in crystal is bigger than atmospheric pressure.

\* *Hygroscopic*: crystals gain water and are used to dry or dehumidify

Vapor pressure in crystal is smaller than atmospheric pressure.

\* *deliquescent crystals* gain enough water to dissolve and form a solution.

4. Heterogeneous Systems: visible particles when mixed

a. Suspensions: Particles that separate upon standing and have layers.

The particles are greater that 100 nm.

b. Colloids: Particles that are 1 nm to 100 nm and stay intermixed. These particles are visible with a beam of light. The scattering of light is called the Tyndall Effect. The movement of the particles is called Brownian

motion. Colloids are often milky when concentrated.

\*Emulsion: a liquid/liquid colloid (oil and water using egg or soap)

**Chapter 16**

1. Solution Formation

a. Factors that effect the speed of dissolving.

\*stirring (agitation)

\*temperature

\*surface area/particle size

\*miscible vs. immiscible (will they dissolve in each other based on bonds)

2. Types of Solutions

a. Unsaturated: It can dissolve more solute at the given temperature.

Test: put in any solute and see if it dissolves.

b. Saturated: It contains the maximum amount of solute for a solvent at a

given temperature. Test: put in any solute and see if it does not dissolve.

(Solute will sit on the bottom or top of the solvent and do nothing.)

c. Supersaturated: It contains more solute than it should at a given

temperature. Test: If put any solute in the solution or disturb it in any

way, the whole solution crystallizes.

3. Temperatures and solubility’s: Trends

\* Increase temperature and there will be an increase in solubility of most solids.

\* Increase temperature and there will be an decrease in solubility of most gases.

\* Increase pressure and there will be an increase in solubility of most gases.

4. Henry’s Laws: at a given temperature, the solubility of a gas in a liquid is

directly proportional to the pressure of the gas above the liquid. (When one goes up, the other goes up: ↓P =↓S and ↑P =↑S)

**S1 = S2**

**P1 P2**

*Example:* If the solubility if a CO2 in water is .77 g/l at 3.5 atmospheres of

Pressure, what is the solubility in g/l at 1.0 atmospheres of pressure when

the temperature is held constant?

5. Concentrations of Solutions

*Qualitative*: a description (senses)

\*concentrated: too much solute or not enough solvent

\*dilute: not enough solute or too much solvent

(unusually amount of solvent is altered.)

*Quantitative*: a numerical measurement

a. **Molarity**: the number of moles of solute dissolves in a liter of solution.

**Molarity = moles**

**Liters**

Example: A solution contains .90 grams of NaCl in 100 ml of solution.

What is the molarity of the solution?

Example: Calculate the molarity of a solution composed of 103 grams of

Ca(OH)2 dissolved in 700 ml of solution.

Example: I need 250 ml of a .10 M sucrose solution for lab. How many

moles of sucrose is needed? How many grams of C12H22O11 do I need? How do I make this solution?

Example: Calculate the number of grams needed to prepare 175 ml

Of 1.20 M NaCl solution. (Hint: Find moles of NaCl first.)

b. **Dilutions:** When solvent is added to lower the concentration of a

solution. The number of moles of solute before and after a dilution remains the same.

Derivation of formula:

**M1 \* V1 = M2 \* V2**

*Example:* How many milliliter of a 6M HCl stock solution are needed to

make 250 ml of 1 M HCl. How do you make this solution? How does

making a solution differ from doing a dilution?

*Example:* You have a stock solution of 4.0 M KNO3. How would your

prepare 750 ml of 3.0 x 10-1 M solution?

c. Percent Solutions

\*If two liquids then: Percent by Volume

% by volume = volume of solute x 100 (same units for volumes)

solution volume

\*If a solid in a liquid then: Percent by Mass

% by mass = mass of solute (g) x 100

solvent volume (ml)

Example: Calculate the percent by mass of a solution made from

dissolving 56.0 grams of C6H12O6 in 2.0 L of water.

Example: A bottle of hydrogen peroxide antiseptic is labeled 3% by

volume. How many ml of H2O2 are in a 500 ml bottle of this solution?

d. **Colligative Properties of Solutions:** this explains how the physical

properties of a solvent change when a solute is added. The more particles,

the bigger the change in the properties. (this includes ions too)

\* *Decrease in the vapor pressure of a solvent*: (It is harder for the solvent

to evaporate because the solute particles are in the way.)

Example: NaCl (s) → Na +(aq) + Cl -(aq) or MgBr2 (s) → Mg +2(aq) + 2Br -(aq)

Which solute lowers the vapor pressure more?

\* *Boiling Point elevation of a solvent*: More kinetic energy (or a higher

temperature) is needed for the solute particles to push the solute particles

aside, overcome atmospheric pressure, and then finally escape or boil.

\* *Freezing Point Depression:* To freeze, the solvent particles must lose

kinetic energy (decrease temperature) and form bonds between solvent

particles. If solute particles are in the way, addition energy is lost to push

the solute out of the way. This causes an decrease in temperature.

e. Calculations of colligative properties.

\* All colligative property calculations use a quantitative

concentration calculation called Molality. Molality is

defined as the moles of solute dissolved per kilogram

of solvent.

**Molality or *m =* moles of solute (m)**

**kilogram of solvent (kg**)

Example: Fine the molality of 3.5 moles dissolve in 1500 grams of water.

Example: How many grams of KI must be dissolved in 500.0

grams of water to produce a .060 *m* solution?

f. **Boiling Point Elevation Problems:** This involves two basic steps.

#1 The change in the boiling = the boiling point x Molality of the

point of the solvent constant of the solvent solution

(pg 495 of book)

**∆Tb = Kb x *m***

#2 The new boiling = (∆Tb x the # of ions in solution) + normal boiling

point of the solvent point

Example: What is the boiling point of a 1.50 *m* NaCl aqueous solution?

Example: What is the boiling point of a solution that contains 1.25 moles

of CaCl2 in 1400 grams of water?

g. **Freezing Point Depression problems**: This involves two basic steps.

#1 The change in the freezing = the freezing point x Molality of the

point of the solvent constant of the solvent solution

(pg 494 of book)

**∆Tf = Kf x *m***

#2 The new freezing = normal freezing- (**∆Tf** x the # of ions in solution)

point of the solvent point

Example: Calculate the freezing point of 200. grams of ethanol C2 H5OH added to 800 grams of water.