Chapter 11 - Properties of Solutions

Section 11.1 - Solution Composition

Solute - substance which is dissolved (if liquid-liquid, the one you have less of is considered the solute)
Solvent - substance which is doing the dissolving.
Molarity (M) - the number of moles of solute per liter of solution.

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{Liters of solution}}$$

Mass percent (weight percent) - percent by mass of solute in the solution.

$$\text{Mass percent} = \left(\frac{\text{mass of solute}}{\text{mass of solution}} \right) \times 100$$

Mole fraction - the ratio of the number of moles of a given component to the total number of moles of solution.

$$\text{Mole fraction of component A} = c_A = \frac{n_A}{n_A + n_B}$$

Molality - the number of moles of solute per kilogram of solvent.

$$\text{Molality} = \frac{\text{moles of solute}}{\text{kilogram of solvent}}$$

Normality - the number of equivalents per liter of solution.

Section 11.2 - Energies of Solution Formation

Three steps in creating a liquid solution
        Step 1 - break up the solute into individual components (expanding the solute).
        Step 2 - overcoming intermolecular forces in the solvent to make room for the solute (expanding the solvent).
        Step 3 - allowing the solute and solvent to interact to form the solution.

Enthalpy (heat) of solution (ΔH_{soln}) - sum of the energies it takes to dissolve a substance.

$$\Delta H_{soln} = \Delta H_1 + \Delta H_2 + \Delta H_3$$

Enthalpy (heat) of hydration (ΔH_{hyd}) - represents the enthalpy change associated with the dispersal of a gaseous solute in water.

        Processes which require large amounts of energy tend not to occur. **See Table 11.3**

Section 11.3 - Factors Affecting Solubility

Since it is the molecular structure that determines polarity, there should be a definite connection between structure and solubility.
Structure effects: determines what kind of substance will be able to dissolve the sample.
        hydrophobic - water-fearing
        hydrophilic - water-loving

While pressure has little effect on the solubilities of solids or liquids, it does significantly increase the solubility of a gas.
Pressure effects: gas solubility increases with increased pressure.

Henry's Law: the amount of a gas dissolved in a solution is directly proportional to the pressure of the gas above the solution.

$$P = kC$$

P represents the partial pressure of the gas, **C** represents the concentration of the dissolved gas, **k** is a constant of the particular solution.

Temperature effects - the solubility of most solids increases with temperature.

Section 11.4 - The Vapor Pressures of Solutions
Nonvolatile - does not readily form a vapor.
Nonvolatile solutes lowers the vapor pressure of a solvent. The nonvolatile solute decreases the number of solvent molecules per unit
   volume. Thus it lowers the number of solvent molecules at the surface, and it should proportionately lower the escaping
   tendency of the solvent molecules.
Raoult's Law:

$$P_{soln} = \chi_{solvent} \, P^o_{solvent}$$

Where P_{soln} is the observed vapor pressure of the solution, $\chi_{solvent}$ is the mole fraction of solvent, and $P^o_{solvent}$ is the vapor pressure of the pure solvent.

If both components of the solution are **volatile** Raoults's law can be modified:

$$P_{TOTAL} = P_A + P_B = \chi_A P^o_A + \chi_B P^o_B$$

Raoult's law works only for **ideal solutions** (like ideal gases). This ideal behavior is often observed when the solute-solute,
   solvent-solvent and solute-solvent interactions are very similar.
If a **lower vapor pressure** occurs than was predicted; there will be a **negative deviation** from Raoult's law.
If a **higher vapor pressure** occurs than was predicted; there will be a **positive deviation** from Raoult's law.
                    See figure 11.13 on page 525

Section 11.5 - Boiling-Point elevation and Freezing-Point Depression
Colligative properties - depend on the number, and not the identity of the solute particles in an ideal solution. ex. - **freezing-point
   depression, boiling-point elevation and osmotic pressure**
Boiling-point elevation - a nonvolatile solute elevates the boiling point of the solvent.

$$\Delta T = K_b m_{solute}$$

Where ΔT is the **boiling-point elevation**, or the difference between the boiling point of the solution and that of the pure
   solvent, K_b is a constant that is characteristic of the solvent is called the **molal boiling-point elevation constant** and m_{solute} is
   the **molality** of the solute in the solution.

Freezing-point depression - the water in the solution has a lower vapor pressure than that of pure ice.

$$\Delta T = K_f m_{solute}$$

Where ΔT is the **freezing-point depression**, or the difference between the freezing point of the pure solvent and that of the
   solution, K_f is a constant that is characteristic of a particular solvent is called the **molal freezing-point depression constant**
   and m_{solute} is the **molality** of the solute in the solution.
                    See Table 11.5 on page 528 - for constants

Section 11.6 - Osmotic Pressure

Osmosis - the flow of solvent into the solution through the **semipermeable membrane**

Osmotic pressure - different liquid levels leads to a greater (or lesser) hydrostatic pressure on the solution than on the pure solvent (**see fig 11.16 on page 531**). The minimum pressure that stops the osmosis is equal to the osmotic pressure of the solution.

$$\pi = MRT$$

Where π is the **osmotic pressure** in atmospheres, **M** is the **molarity** of the solution **R** is the **gas law constant** (0.08206 L atm/K atm), and **T** is the **Kelvin temperature**

Section 11.7 - Colligative Properties of Electrolyte Solutions

van't Hoff factor (i) - the relationship between the moles of solute dissolved and the moles of particles in solution

$$i = \frac{\text{moles of particles in solution}}{\text{moles of solute dissolved}}$$

Ion pairing - a phenomenon occurring in solution when oppositely charged ions aggregate and behave as a single particle. This will cause changes in the freezing and boiling points.

$$\Delta T = i \, m \, K$$

Where **K** represents the **freezing-point depression** or **boiling-point elevation constant** for the solvent.

This will also cause changes in the **osmotic pressure**.

$$\pi = i \, MRT$$

Section 11.8 - Colloids

Tyndall effect - the scattering of light by particles and is used to distinguish between a suspension and a true solution.

Colloid (colloidal dispersion) - suspension of tiny particles in some medium. **Table 11.7 on page 539** Particles do not precipitate out because of electrostatic repulsion.

Coagulation - the destruction of a colloid, usually can be accomplished either by heating or by adding an electrolyte.