SOLUTIONS

Solutions are homogeneous mixtures in which the physical properties are dependent on the concentration of the solute and the strengths of all interactions among the particles of the solutes and the solvent (EK.2.A.3)

Solute - the substance being dissolved.

Solvent - the substance doing the dissolving (usually present in the greatest amount).

HOMOGENEOUS MIXTURE (SOLUTION)
- Macroscopic properties do not vary throughout the sample.

Types of Solution:

Solute	Solvent	Solution Appearance	Example
Gas	Gas	Gas	Air
Gas	Liquid	Liquid	Soda (CO_2 in water)
Liquid	Liquid	Liquid	Vinegar (CH_3COOH in water)
Solid	Liquid	Liquid	Saline solution (NaCl in H_2O)
Solid	Solid	Solid	Brass (Zn in Cu) - an alloy

Liquid Solutions:
- Exhibit several general properties:
    1. The components cannot be separated by using filter paper
    2. There are no components big enough to scatter light
    3. The components can be separated using processes that rely on intermolecular forces between the components.

Separation of Solution components:
Chromatography
- o can be paper or column chromatography
- o separates the components based on their affinity for the stationary phase and the mobile phase. (ie: takes advantage of the intermolecular interactions between all components).
Distillation
- o uses the differences in the strength of intermolecular forces between the components and among the components. These affect the vapor pressures of the components in the mixture.

Solution Concentration and Preparation of Solutions:

(A) **Molarity (M) =** <u>**moles of solute**</u>
                                liters of solution

- Since molarity is dependent on volume, which changes slightly with temperature, it is dependent on temperature.
- Ex's: Calculate the molarity of
  a) 3.00 mole of KI in 2.39 L of solution
  b) 0.522 g HCl in 0.592 L of solution.
  c) How many grams of potassium hydroxide are needed to prepare 10.0 mL of 0.100 M KOH?
  d) A bottle of aqueous formic acid is 90.0% HCOOH by mass and has a density of 1.20 g/mL. Calculate the molarity of the solution.

(B) Dilution of solutions:

$$M_1 \times V_1 = M_2 \times V_2 \qquad \text{(M=molarity \quad V=volume)}$$

- Ex.'s
  a) How many mL of a 12 M solution of HCl are required to make 250 mL of a 0.10 M solution?
  b) How many mL of 5.0 M $K_2Cr_2O_7$ must be diluted in order to prepare 250 mL of 0.10 molar solution?

(C) Preparation of solutions:
  - You should know and be able to describe how to prepare a solution of a particular molarity either from a solid solute or from a solution of known molarity.
  Important points to remember:

  - Always use the most accurate glassware available. This will usually be a volumetric flask and possibly a measuring cylinder, buret, or pipet.
  - The volume is total volume (solute and solvent together). You will need to mix the solute and solvent in a volumetric flask then add water to get the correct total volume.
  - You cannot assume that volumes are additive unless it is specified (50mL + 50 mL does not always equal 100 mL).
  - Distilled or deionized water should always be used.
  - If a strong acid is being diluted always add the acid to the water with plenty of swirling.

Mole Fraction (X) = moles of component A / Total moles   ($n_A/ n_A + n_B \ldots$)

- note there are no units
- the sum of all the mole fractions for all the components in a solution is 1
- **mole percent = mole fraction x 100**

Solution Formation:
- May be exothermic or endothermic depending on the relative strengths of intermolecular forces between all of the components before and after the dissolution process (EK.2.A.2)

The main factors affecting the dissolution of a solute in a solvent are:
1. solute to solute interactions.
2. solvent-solvent interactions.
3. solvent-solute interactions.

Dissolution (formation of a homogeneous mixture) will be favored when the first two are relatively small and the third is relatively large.

Like dissolves like

- in general, if two substances have similar intermolecular forces, the solute should have high solubility in the solvent.

Enthalpy of Solution (heat of solution).

- Consider solution formation as three hypothetical steps, each with an enthalpy change associated with it.
  Step 1: Separating solvent molecules ΔH_1
  Step 2: Separating solute molecules ΔH_2
  Step 3: Allowing the solute and solvent to interact to form the solution ΔH_3
  $\Delta H_{solution} = \Delta H_1 + \Delta H_2 + \Delta H_3$

1. **Ideal solutions** $\Delta H_{solution} = 0$ $\Delta V_{solution} = 0$
   - All intermolecular forces are of comparable strength.
   - No net energy change on solution formation.
   - Volume of solution is the sum of the volumes of the solute and solvent.
   - Mixtures of ideal gases
   - Some liquids come close to forming ideal solutions. Ex: toluene in benzene.

2. $\Delta H_3 > \Delta H_1 + \Delta H_2$ ($\Delta H_{solution} < 0$ $\Delta V_{solution} < 0$)
   - IE: intermolecular forces between solute and solvent particles are greater than other intermolecular forces.
   - Solution formation is exothermic
   - Volume of the solution is less than the sum of the solute and solvent volumes.
   - Nonideal.

3. $\Delta H_3 < \Delta H_1 + \Delta H_2$ ($\Delta H_{solution} > 0$)
   - IE: intermolecular forces between solute and solvent particles are weaker than other intermolecular forces.
   - Solution formation is endothermic.
   - May occur due to an increase in entropy (disorder).
   - Nonideal.

4. $\Delta H_3 < << \Delta H_1 + \Delta H_2$
   - The intermolecular forces between solute and solvent are so much smaller than the sum of the other intermolecular forces that the solution may not form - the two remain as separate (heterogeneous) phases.

Ionic Compounds:
   - whether they dissolve in water depends on whether the ion-dipole interactions between ions and water molecules are stronger than the interionic interactions.
   - generally, when dissolved in water, the components are separated and dispersed. The presence of ions can be detected using conductivity measurements.

Electrolytes
- ionize in solution.
- **Strong electrolytes** - solutes that exist in solution almost exclusively as ions. Most ionic compounds, and a few molecular compounds (mainly strong acids such as HCl.
- **Weak electrolytes** - solutes that are only partially ionized in solution. Some molecular compounds, carboxylic acids, and amines.
- **Nonelectrolytes** - solutes that exist in solution almost exclusively as molecules. Most molecular and organic compounds.

Key Terms:

Miscible liquids will mix together in all proportions.

Immiscible liquids do not mix and will form heterogeneous mixtures.

Solvation - the process of dissolving a solute in a solvent

Hydration - the same as solvation, but the term is used when the solvent is water.

Hydrophobic - (water hating) is often used to describe nonpolar substances.

Hydrophilic - (water loving) is often used to describe polar substances.

Factors Affecting Solubility of Solutions:

Equilibrium:

- Dynamic equilibrium occurs when the rate of crystallization is equal to the rate of dissolving.
- **Saturated** solutions have reached dynamic equilibrium. (The solution contains all the dissolved solute it can hold at that temperature).
- The **solubility** of a solute is the concentration of the solution at equilibrium.
- **Unsaturated** solutions contain less solute than can be held at equilibrium.
- **Supersaturated** solutions contain more solute than can be held at equilibrium.

Temperature Effects:

- Solubility changes with temperature.
- If a dissolution process is exothermic, the solubility of a solute will decrease with an increase in temperature.
- If a dissolution process is endothermic, the solubility of the solute will increase with an increase in temperature.
- Most ionic compounds have aqueous solubilities which increase with increased temperature.
- Solubility curves graph solubility v's temperature.
- Supersaturated solutions may be formed by dissolving the solute at a higher temperature then cooling the solution slowly. (Supersaturated solutions are very unstable).

For Gases:

- Most gases are less soluble in water as temperature increases but often become more soluble in organic solvents as temperature increases.

Pressure Effects:

- Changes in pressure have very little effect on the solubilities of solids or liquids in liquid solutions.
- Gas solubilities in all solvents are affected by changes in pressure.
- **Henry's Law $S = k\,P$**   (S=solubility, k=constant depending on the particular gas and temp P=pressure of the gas).

HETEROGENEOUS MIXTURES:
- The macroscopic properties depend on the location in the mixture.

- The distinction between homogeneous and heterogeneous may depend on the scale you are looking at.

Colloids:

Solution - Homogeneous mixture with particles less than 1 nm.

Suspension - heterogeneous mixture where solute is temporarily dispersed through the solvent. Particle sizes are > 1000 nm.

Colloid - dispersion of particles with sizes from 1 nm to 1000 nm. Colloidal particles cannot generally be seen through an ordinary microscope. ie: colloids may seem heterogeneous if you are looking on the scale of micrometers but may appear homogeneous if you are looking on the scale of centimeters.

Properties of colloids:
- Often appear milky or cloudy
- **Tyndall effect** - a beam of light passed through a colloid will scatter.
- **Coagulation** (destruction of a colloid) can be attained by heating the solution or by adding an electrolyte.

Types of Colloids

Dispersed Phase	Dispersion Medium	Colloid Type	Examples
Solid	Liquid	Sol	Paint, clays, gelatin
Liquid	Liquid	Emulsion	Milk, mayonnaise,
Gas	Liquid	Foam	Whipped cream, soap suds, meringue
Solid	Gas	Aerosol	Smoke, dust, airborne bacteria
Liquid	Gas	Aerosol	Fog, hair sprays
Solid	Solid	Solid sol	Ruby glass, black diamond
Liquid	Solid	Solid emulsion	Butter, pearl, opal
Gas	Solid	Solid foam	Marshmallow, pumice

Examples: Predict whether the following are likely to be solutions or heterogeneous mixtures. A) methanol and water B) pentane and octane C) sodium chloride and carbon tetrachloride D) 1-decanol and water.