

Chapter 13: Properties of Solutions

13.1 The Solution Process

- solution formed when one substance disperses uniformly throughout another
- **solvation** – interaction between solute and solvent molecules
- **hydration** – solvation when solvent is water

13.1.1 Energy Changes and Solution Formation

- overall enthalpy change in formation of a solution
 - $\Delta H_{\text{soln}} = \Delta H_1 + \Delta H_2 + \Delta H_3$
 - ΔH_1 = separation of solute molecules
 - ΔH_2 = separation of solvent molecules
 - ΔH_3 = separation formation of solute - solvent interactions
- separation of solute particles is endothermic
- separation of solvent is endothermic
- third is exothermic
- formation of solution can be either exothermic or endothermic
- exothermic processes are spontaneous
- solution will not form if enthalpy is too endothermic
- H_3 has to be comparable to $H_1 + H_2$
 - Ionic substances cannot dissolve in nonpolar liquids
 - Polar liquids do not form solutions with nonpolar liquids

13.1.2 Solution Formation, Spontaneity, and Disorder

- two nonpolar substances dissolve in one another
- attractive forces = London dispersion forces
- two factors in processes that are spontaneous: energy and disorder
- processes in which the energy content of the system decreases tend to occur spontaneously
 - exothermic
- processes in which the disorder of the system increases tend to occur spontaneously
- solutions will form unless solute-solute or solvent-solvent interactions too strong relative to solute-solvent interactions

13.1.3 Solution Formation and Chemical Reactions

- distinguish between physical process of solution formation from chemical process that leads to a solution

13.2 Ways of Expressing Concentration

- dilute and concentrated used to describe solution qualitatively
- mass percentage of component in solution:
 - $\text{mass\% of component} = \frac{\text{mass of component in soln}}{\text{total mass of soln}} \times 100$
- very dilute solutions expressed in parts per million (ppm)
 - $\text{ppm of component} = \frac{\text{mass of component in soln}}{\text{total mass of soln}} \times 10^6$
- 1ppm = 1g solute for each (10^6) grams of solution or 1mg solute per kg solution

- 1ppm = 1 mg solute/L solution
- 1 ppb = 1 g of solute/ 10^9 grams of solution or 1 μg solute/ L of solution

13.2.1 Mole Fraction, Molarity, and Molality

- mole fraction of component = $\frac{\text{moles of component}}{\text{total moles of all components}}$
- sum of mole fractions of all components of solution must equal one
- Molarity = $\frac{\text{moles solute}}{\text{liters soln}}$
- molality = $\frac{\text{moles solute}}{\text{kilograms of solvent}}$
- molality goes not vary with temperature
- molarity changes with temperature because of expansion and contraction of solution

13.3 Saturated Solutions and Solubility

- **crystallization** – reverse process of solution
- dynamic equilibrium – when equilibrium exists between process of solution and crystallization
 - solute said to be saturated
- **solubility** – amount of solute needed to saturate a solution
- **unsaturated** – when there isn't enough solute to saturate a solution
- **supersaturated** – when there is more solute than needed to saturate a solution
- for most salts crystallization of excess solute is exothermic

13.4 Factors Affecting Solubility

13.4.1 Solute-Solvent Interactions

- solubility increases with increasing molar mass
- London dispersion forces increase with increasing size and mass of gas molecules
- **Miscible** – pairs of liquids that mix in all proportions
- **Immiscible** – opposite of miscible
- Hydrogen-bonding interactions between solute and solvent leads to high solubility
- Substances with similar intermolecular attractive forces tend to be soluble in one another
- “like dissolves like”

13.4.2 Pressure Effects

- solubility of a gas in any solvent increases as pressure of gas over solvent increases
- relationship between pressure and solubility: Henry's Law:
 - $C_g = kP_g$
 - C_g solubility of gas in solution phase (usually expressed as molarity), P_g partial pressure of gas over solution, k is proportionality constant (Henry's Law constant)
 - Henry's law constant different fore each solute-solvent pair, and temperature

13.4.3 Temperature Effects

- solubility of most solid solutes in water increases as temperature of solution increase
- solubility of gases in water decreases with increasing temperature
- decreases solubility of O_2 in water as temperature increases in one the effects of thermal pollution

13.5 Colligative Properties

- **colligative properties** – physical properties that depend on quantity

13.5.1 Lowering the Vapor Pressure

- vapor pressure over pure solvent higher than that over solution
- vapor pressure needed to obtain equilibrium of pure solvent higher than that of solution

13.5.2 Raoult's Law

- Raoult's law: $P_A = X_A P_A^\circ$
- P_A = vapor pressure of solution, X_A = mole fraction of solvent, P_A° = vapor pressure of the pure solvent
- **Ideal solution** – solution that obeys Raoult's law
 - Solute concentration is low, solute and solvent have similar molecular sizes and similar types of intermolecular attractions

13.5.3 Boiling-Point Elevation

- normal boiling point of pure liquid is the temperature at which pressure is 1 atm
- addition of a nonvolatile solute lowers vapor pressure of solution
- $\Delta T_b = K_b m$
- K_b = molal boiling-point-elevation constant
 - Depends only on solvent
 - boiling point elevation proportional to number of solute particles present in given quantity of solution

13.5.4 Freezing-Point Depression

- freezing point of solution is temperature at which the first crystals of pure solvent form in equilibrium
- freezing point of solution lower than pure liquid
- freezing point directly proportional to the molality of the solute:
 - $\Delta T_f = K_f m$
 - K_f = molal freezing-point-depression constant

13.5.5 Osmosis

- **semipermeable** – membranes that allow passage of some molecules and not others
- **osmosis** – the net movement of solvent molecules from the less concentrated solution to the more concentrated solution
- net movement of solvent always toward the solution with the higher solute concentration
- osmotic pressure – pressure needed to prevent osmosis, π
- $\pi = \left(\frac{n}{V}\right)RT = MRT$
- M = molarity of solution
- if solutions identical osmosis will not occur and said to be **isotonic**
- if one solution lower osmotic pressure = **hypotonic**, the solution that has higher osmotic pressure = **hypertonic**
- **crenation** = when cells shrivel up from the loss of water
- **hemolysis** = when cells rupture due to too much water

13.5.6 Determination of Molar Mass

- see examples 13.10, 13.11 in book
- colligative properties can be used to find molar mass

13.6 Colloids

- **colloidal dispersions, colloids** – intermediate types of dispersions or suspensions
- intermediate solutions between solutions and heterogeneous mixtures
- colloids can be gases, liquids, or solids

- colloid particles have size between 10 - 2000Å
- **tyndall effect** – scattering of light by colloids

13.6.1 Hydrophilic and Hydrophobic Colloids

- **hydrophilic** – colloids in which the dispersion medium is water
- **hydrophobic** – colloids not dispersed in water
- hydrophobic colloids have to be stabilized before being put in water
 - natural lack of affinity for water causes separation
 - can be stabilized by the adsorption of ions on the surface
 - adsorped ions interact with water
 - can also be stabilized by presence of other hydrophilic groups on surface

13.6.2 Removal of Colloidal Particles

- **coagulation** – enlarging colloidal particles
 - heating or adding an electrolyte to mixture
 - heating increases number of collisions and particles stick together increasing their size
 - electrolytes causes neutralization of the surface charges of the particles which remove the electrostatic repulsion
- **dialysis** – use of semipermeable membranes to filter out colloidal particles