

AP Chemistry

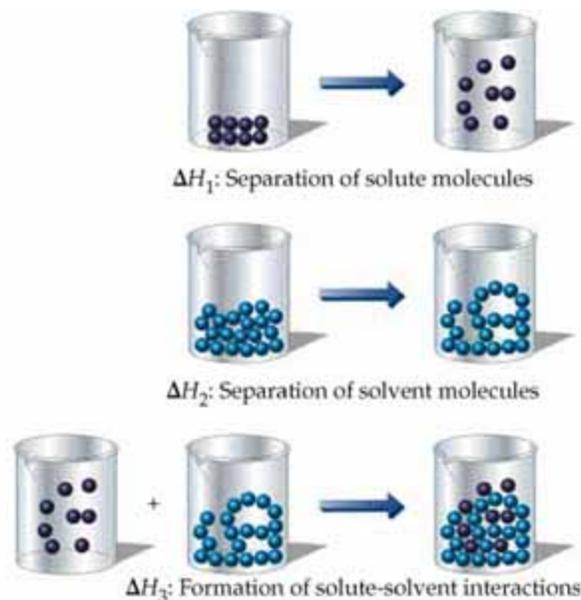
Chapter 13 Outline

- The Solution Process
 - Solution formation occurs when one substance disperses uniformly throughout another
 - Due to intermolecular attractions between solute and solvent particles
 - Strictly physical changes: solute can be recovered unchanged from the solution
 - Solvation = interactions between solvent and solute
 - Hydration = when solvent is water
 - Solutions form when solvent-solute interactions overcome the attractions of the solute for other solute particles

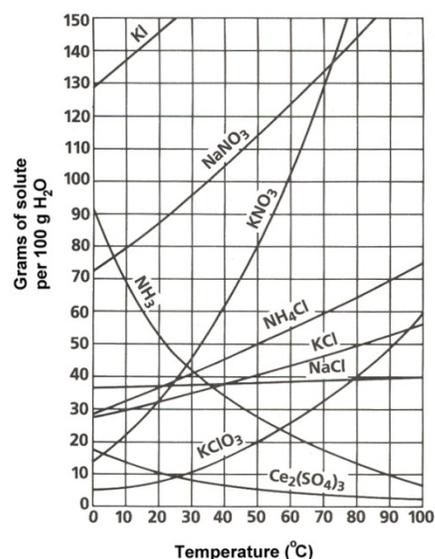
- Formation of a solution can be thought of as a three step process $\Delta H_{\text{soln}} = \Delta H_1 + \Delta H_2 + \Delta H_3$

- ΔH_1 = separation of solute particles (endothermic!)
- ΔH_2 = separation of solvent particles (endothermic!)
- ΔH_3 = formation of solute-solvent interactions (exothermic!)

- Formation of a solution can be either exothermic or endothermic!
 - Processes that are exothermic tend to proceed spontaneously.
 - A solution will not form if ΔH_{soln} is too endothermic



- Saturated Solutions and Solubility
 - As a solid solute dissolves, its concentration in the solvent increases, which in turn raises the chances of colliding with the surface of the solid.
 - The solute particle may reattach to the solid = crystallization
 - Solubility equilibrium may result (see section 11.5)
 - **Saturated solution** = holds maximum amount of solute for that temperature; in equilibrium with undissolved solute
 - **Unsaturated solution** = holds less than maximum amount of solute; able to dissolve more solute
 - Supersaturated solution = contains a greater amount of solute than a saturated solution; unstable



- Solubility of a solution = the amount of solute needed to form a saturated solution in a given quantity of solvent at a certain temperature
 - For most solids, solubility **increases** as solution temperature increases.
 - There are exceptions!
 - For most gases, the solubility decreases with increasing solution temperature.
- Factors Affecting Solubility
- In general, the stronger the attractions between solvent and solute molecules, the greater the solubility.
 - “Like dissolves like”: substances with similar intermolecular attractive forces tend to be soluble in one another
 - Polar liquids tend to dissolve readily in polar solvents.
 - Why? Favorable [dipole-dipole attractions](#)
 - Miscible = pairs of liquids that mix in all proportions
 - Non-polar liquids tend to be insoluble in polar liquids.
 - Immiscible = liquids that do not dissolve in each other
 - [Alcohols](#)
 - –OH group is polar and can form hydrogen bonds with water and
 - As hydrocarbon tail on alcohol molecule gets longer, alcohol becomes less soluble in water and more soluble in hexane
 - Ionic substances tend to dissolve in polar substances.
 - [Ion-dipole attractions](#)
 - Covalent substances tend to dissolve in nonpolar solvents.
 - Induced dipoles ([London forces](#)) are very weak, and if the solvent can do any other type of attraction, it’s not energetically favorable to interact with the nonpolar solvent
 - Network solids have very strong bonding forces in the solid, so they tend to not dissolve in either polar OR nonpolar liquids.
- A solution will typically form, unless solute-solute and/or solvent-solvent attractions are stronger than the solute-solvent interactions.
- Entropy (to be revisited in a later chapter): Processes occurring at constant temperature which result in an increase in randomness of the arrangement of the particles tend to occur spontaneously.
- Pressure Effects (gases)
- The solubility of any gas is increased as the pressure over the solvent increases.
 - The solubility of the gas increases in direct proportion to its partial pressure above the solution.
 - [Henry’s Law](#): $S_g = kP_g$
 - S_g = solubility of gas in solution phase (usually, molarity)
 - P_g = partial pressure of gas over solution
 - k = Henry’s law constant; different for each solute-solvent pair; varies with temperature

- Observe Henry's Law when you open a bottle of a carbonated beverage!
 - When opened, P_{CO_2} decreases, so CO_2 becomes less soluble and bubbles out of the beverage

Ways of Expressing Solution Concentration This section is absolutely critical!

- Qualitative descriptions
 - Dilute = solution with relatively small amount of solute
 - Concentrated = solution with relatively high amount of solute
- Quantitative descriptions
 - **Mass percentage** (no longer tested on AP exam, but very important in everyday use!)

$$\text{Mass}\% = \frac{\text{mass of component in solution}}{\text{total mass of solution}} \times 100$$

- **Part per million**—useful for very dilute solutions
 - 1 ppm = 1 mg per kilogram of solution (“mg per kg”)

$$\text{ppm of component} = \frac{\text{mass of component in solution}}{\text{total mass of solution}} \times 10^6$$

- **Parts per billion** = 1 μg per kilogram of solution
- **Mole fraction**, symbolized X with a subscript to identify the substance; particularly useful with gases
 - Sum of mole fractions of all components must add up to 1

$$\text{Mole fraction of component} = \frac{\text{moles of component}}{\text{total moles of all components}}$$

- **Molarity**—the chemist's #1 choice for concentration unit!
 - symbol M
 - Molarity of solution changes with temperature because volume of solution can change with temperature

$$M = \frac{\text{moles solute}}{\text{liters solution}}$$

- **Molality** (no longer tested on the AP Chemistry exam)
 - symbol m
 - Molality of solution does NOT change with temperature because masses do not change with temperature

$$m = \frac{\text{moles solute}}{\text{kg of solvent}}$$

You need to be able to convert between different concentration units.

- Colligative Properties (This material is no longer tested on the AP Chemistry exam)
 - Properties of solutions that depend on the NUMBER of solute particles (not the type)
 - **Lowering the Vapor Pressure**
 - Vapor pressure = the pressure exerted by the vapor above its own liquid in a closed container at equilibrium
 - Nonvolatile = a substance with no measurable vapor pressure
 - Volatile = exhibits a vapor pressure
 - Adding a nonvolatile solute to a solvent always lowers the vapor pressure of the solvent.
 - Raoult's Law = the partial pressure exerted by solvent vapor above a solution, P_A , equals the product of the mole fraction of A, X_A , times the vapor pressure of the pure solvent, P_A°

$$P_A = X_A P_A^\circ$$

- Depends on the total concentration of dissolved particles! Ions OR molecules
- Ideal solutions obey Raoult's law.
 - Real solutions best approach ideal behavior if solution concentration is low and when solute and solvent particles are similar in size and in types of intermolecular attractions.

- **Boiling Point Elevation**

- The boiling point of the solution is higher than that of the pure liquid.
 - The elevation in boiling point is directly proportional to the number of solute particles

$$\Delta T_b = K_b i m$$

- K_b = molal boiling point elevation constant, depends on solvent; will be given
- Van't Hoff factor i = measure of extent to which electrolytes dissociate
 - Theoretical value of i = number of ions present in formula ($i = 1$ for covalent molecules)
 - The more dilute the solution, the more closely i approaches the theoretical value
 - In concentrated solutions, i is less than the ideal value due to "ion pairing": ions of opposite charge collide and "stick together" briefly, which lowers the number of independent particles in solution

- **Freezing point depression**

- The freezing point of the solution is lower than that of the pure solvent.

$$\Delta T_f = K_f i m$$

- K_f = molal freezing point depression constant; given for each solvent

- **Osmotic Pressure, π**

- In osmosis, the net movement of solvent through a semipermeable membrane is always toward the solution with the higher solute concentration.
- If pressure is placed on the solution with higher concentration, it is possible to stop osmosis of solvent! This pressure is called osmotic pressure.

$$\pi = \left(\frac{n}{V}\right)RT = MRT$$

- Colligative properties are very useful for experimentally determining molar mass.

- **Colloids**—intermediate between solutions and suspensions

- Appear cloudy or opaque, unless very dilute
- Particles are too small to separate by filtration
- Exhibit the **Tyndall effect** (scattering of light)

Phase Diagrams (no longer tested on the AP exam)

Normal boiling point: the temperature at which the substance boils (liquid \rightarrow gas) at 1 atm

Normal freezing point: the temperature at which the substance freezes (liquid \rightarrow gas) at 1 atm

Triple point: the temperature and pressure at which all three phases coexist

Critical point: the temperature and pressure above which a gas cannot be condensed into a liquid

