

# Solutions and Their Properties

Inspired by Paul Groves

## A BLUFFER'S GUIDE

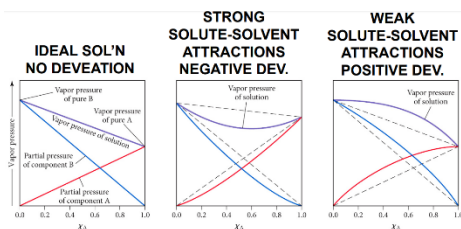
- Solutions are homogenous mixtures.
  - Solute = substance dissolved
  - Solvent = the dissolving medium
- Solutions can be:
  - Unsaturated – less than max solute dissolved at a specific temperature.
  - Saturated – max solute dissolved at a specific temperature.
  - Supersaturated – more than max solute dissolved at a specific temperature.
    - Can be done by heating solvent up, dissolving a lot of solute, then carefully and slowly cooling it back down to the temperature of interest. Unstable, can quickly “crash out” the extra solute if bumped.
- Solutes can be:
  - Electrolytes – conduct electricity when in an aqueous solution
  - Non-electrolytes – do not conduct electricity when in an aqueous solution
- Solids and gases can be soluble or insoluble.
  - Solids increase solubility as T increases. Dissolve better when heated, stirred, ground into small particles.
  - Gases increase solubility as T decreases and P increases. Dissolve better when the solution is cold and the pressure is high.
- Liquids can be miscible or immiscible
- “Like dissolves like”
  - Polar solutes dissolve/mix best in polar solvents
  - Non-polar dissolve/mix best in non-polar
- Concentration can be calculated many different ways. Most common ones:
  - $Molarity = M = \left( \frac{\text{moles of solute}}{\text{Liters of solution}} \right)$
  - $Mass\ Percent = \left( \frac{\text{mass of solute}}{\text{mass of solution}} \right) \times 100$
  - $Mole\ fraction\ of\ A = X_A = \left( \frac{n_A}{n_A + n_B + \dots} \right)$
  - Don't forget that you can use things like density to find numbers you are missing for concentration calculations
  - See your reference sheet for other ways of calculating concentration!

- Heat of solution is the amount of heat absorbed or released when a specific amount of solute dissolves in a solvent.
  - $\Delta H_{\text{solution}} = \Delta H_{\text{solute}} + \Delta H_{\text{solvent}} + \Delta H_{\text{mix}}$
- To make a solution you must:
  - Overcome the attractions between solute particles to separate them.  
 $\Delta H_{\text{solute}} = \text{endothermic}$
  - Overcome some attractions between solvent molecules to make room for the solute.  
 $\Delta H_{\text{solvent}} = \text{endothermic}$
  - Form new attractions between solute particles and solvent molecules.  
 $\Delta H_{\text{mix}} = \text{exothermic}$

Solvent-solute interactions	>	Solvent-solvent and solute-solute interactions	Solution forms
	=		Solution forms
	<		Solution may or may not form, depending on relative disparity

- Factors favoring solution formation:
  - Negative (exothermic)  $\Delta H_{\text{solution}}$
  - Positive  $\Delta S$ , increase in entropy
  - Sometimes an entropy may be negative because the solvent may make “shells” surrounding the solute that results in a decrease in entropy! Often happens with water, referred to as “spheres of hydration”
- A closed container of liquid will have some vapor gas particles above the surface. This exerts a pressure above the liquid.
- ~~Raoult's Law states that the partial pressure of each component in an ideal mixture of liquids is equal to the vapor pressure of the pure compound multiplied by its mole fraction in the mixture.~~
  - ~~$P_{\text{solvent}} = X_{\text{solvent}} P_{\text{solvent}}^{\circ}$~~
  - ~~$P_{\text{total}} = P_A + P_B = X_A P_A^{\circ} + X_B P_B^{\circ}$~~
  - ~~$\Delta P = P_{\text{solvent}}^{\circ} - P_{\text{solution}} = X_{\text{solute}} P_{\text{solvent}}^{\circ}$~~
  - ~~Therefore, presence of a nonvolatile solute lowers the vapor pressure of the solvent.~~

13. Solutions can deviate from Raoult's Law because they are not ideal
- The higher the IMFs in the liquid, the lower the vapor pressure above, because fewer particles will enter the vapor phase. The lower the IMFs in the liquid, the higher the vapor pressure above, because more particles can enter the vapor phase.
  - Negative deviations = lower than predicted vapor pressure, b/c strong IMFs between solute and solvent.
  - Positive deviations = higher than predicted vapor pressure, b/c weak IMFs between solute and solvent.



14. Colligative properties are properties that are affected by how many particles are in a solution, not the identity of the particles.

- Boiling point elevation
- Freezing point depression
- Osmotic pressure

15. If a substance dissolves into multiple ions when in solution it will have a larger effect on the colligative properties

- van't Hoff Factor,  $i$ , the number of particles that a substance breaks apart into.
  - Example  $\text{NaCl} = 2$ ,  $\text{CaCl}_2 = 3$
- $\Delta T = i \cdot K_f \cdot m_{\text{solute}}$
- $\Delta T = i \cdot K_b \cdot m_{\text{solute}}$

16. Separation techniques exploit differences in various properties.

- Filtration – particle size
- Decanting – miscibility
- Distillation – boiling point
- Chromatography – polarity

17. Dilution problems use  $M_1V_1 = M_2V_2$

- Careful because sometimes a problem will ask how much water you need to add to  $V_1$  in order to reach  $V_2$

18. Concentration can be measured using spectrophotometers sometimes.

- Taught in the kinetics chapter
- Beer's Law  $A = \epsilon b C$

19. Net Ionic equations show which particles are actually participating in the reaction.

- Spectator ions are not shown
- Phases should always be included
- If two aqueous substances combine to form an insoluble product, the insoluble product is called a precipitate (ppt)

20. The following ions are considered always soluble.

- Alkali metals
- Ammonium
- Acetate
- Chlorate
- Nitrate
- Perchlorate

21. Particle diagrams are a way to visually show which particles are in various phases or proportions.



22.  $K_{sp}$  is an application of equilibrium in which a solid reactant is dissociating into aqueous products.

- Example:  $\text{A}_2\text{B}(s) \rightarrow 2\text{A}^+(aq) + \text{B}^{2-}(aq)$   
 $K_{sp} = [\text{A}^+]^2[\text{B}^{2-}]$
- Molar solubility "s" can be calculated from  $K_{sp}$  equation. Using the same equation as above:  $K_{sp} = (s)^2(2s)$

23. The common ion effect is an application of equilibrium in which the presence of a common ion to the system being studied can affect the amount of solute that will dissociate.

- Example:  $\text{NaOH}(s) \rightarrow \text{Na}^+(aq) + \text{OH}^-$   
 If you add NaCl it will shift equilibrium to the left, decreasing the solubility of the NaOH