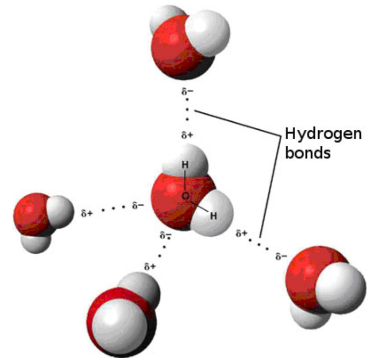
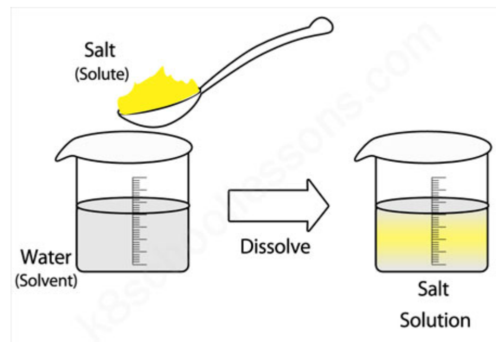
Unit 7 Notes - Aqueous Solutions and Acid/Base Theory

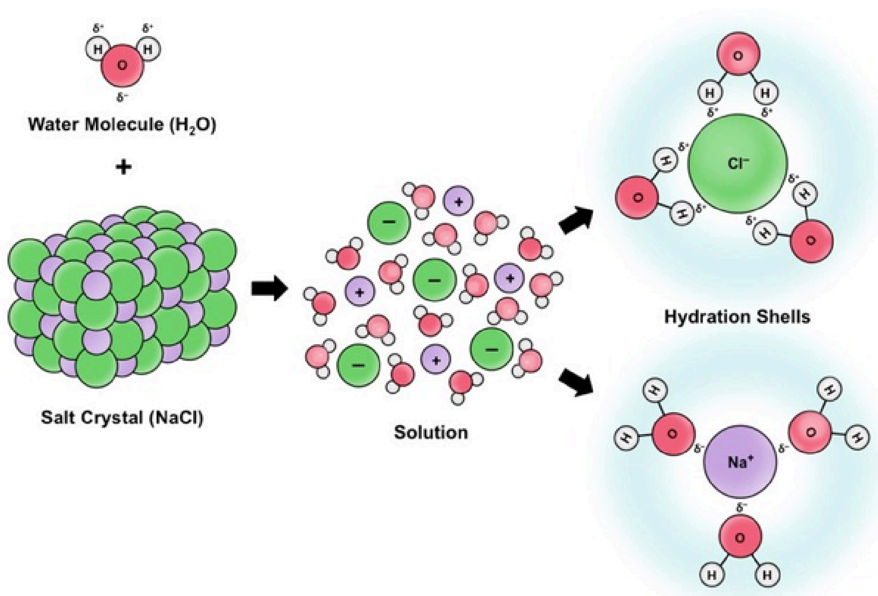
1. Unique properties of water
    - Polarity - The water molecule is very polar. Very strong IMAs between molecules, leading to the other properties below.
    - Cohesion/Adhesion - Attractive forces between water molecules (cohesion) and between water and other substances (adhesion)
    - Capillary action - Water climbs thin tubes called capillaries
    - Surface tension - A “skin” over the surface of water.
    - Heat capacity - Specific heat of water is 4.18 J/g°C, unusually high
    - Density - Water becomes less dense as it freezes
    - Universal solvent - Strong polar bonds means water is a very good solvent



2. Solutions
    A. Composition of a solution
        a. Solvent - “dissolver.” Can be polar/nonpolar, solid/liquid or gas.
        b. Solute - “dissolvee.” Can be polar/nonpolar/ionic. Can be solid/liquid or gas.
        c. Solutions vs suspensions vs colloids - Differ in terms of particle size. Suspensions large particle size, settle over time. Solution small particle size, does not settle.
            i. Tyndall effect - Beam of light or laser reflects off of colloid particles.

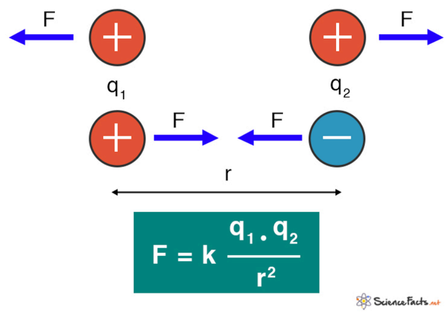


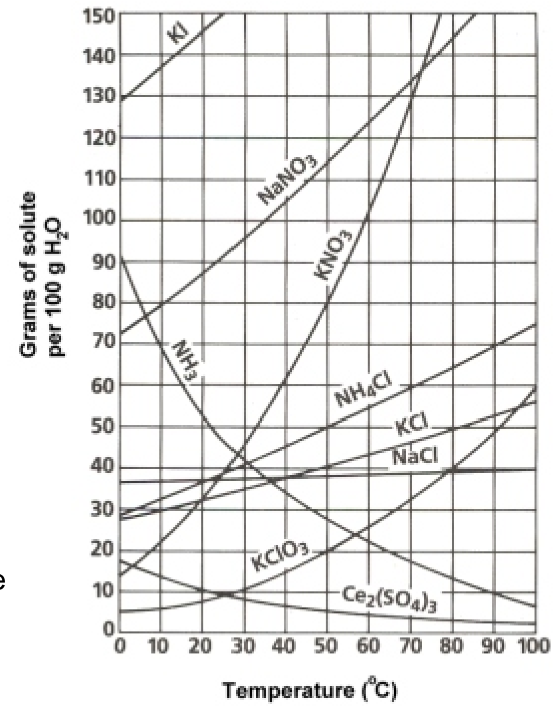
    B. Solvation (dissolution) process
        a. Dissociation - Particles are pulled apart from one another due to their attractions to water molecules
            i. IMA’s & sphere of hydration - positive and negative ends of water molecules (solvent) attract to the opposite charges on the solute, surrounding and dividing them.
            ii. “Like dissolves like” - Solutes with charges (ionic/polar) dissolve in polar solvents. Nonpolar solutes dissolve in nonpolar solvents



C. Solubility

  a. Lattice energy & Coulomb's law



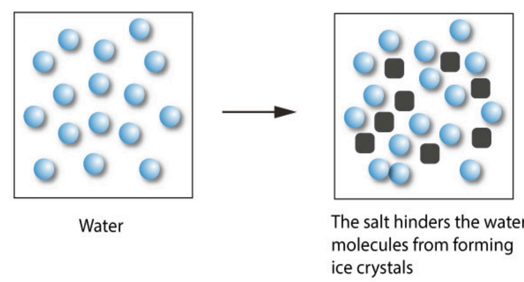


  b. Saturation point - Max amount of solute is dissolved. Less than this is "unsaturated."
  c. Supersaturation - Can manipulate temperature to get above saturation point
  d. Effects on solubility
    i. Temperature - Higher temps allow more solute to dissolve (reverse for gasses)
    ii. Pressure - Higher pressure causes more solubility of gas solutes

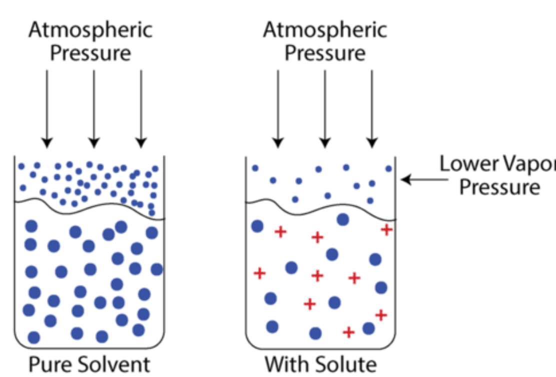
D. Concentration
  a. % by mass/volume - mass solute/mass solution * 100 (or volume)
  b. PPT/PPM/PPB - Parts per thousand, per million, per billion
  c. Molarity (M) - Moles solute / Liter solution
  d. Molality (m) - Moles solute / kg solvent

E. Dilutions
  a. Stock solution - Highly concentration solutions meant for storage. Diluted to the desired concentration
  b. Dilution equation (Rodrigo's Law): $M_1V_1=M_2V_2$

F. Colligative properties
  a. Freezing point depression - Adding solute to a solution lowers the freezing point temperature by disrupting crystal formation
    i. $\Delta T = iK_fm$
      1. K_f for water = 1.86 °C kg/mol
      2. m = molality
      3. i = van't hoff factor. # of dissolved particles in solution



  b. Boiling point elevation - Adding solute to a solution raises the boiling point temperature by making it more difficult for particles to vaporize
    i. $\Delta T = iK_bm$
      1. K_b for water = 0.512 °C kg/mol,
      2. m = molality
      3. i = van't hoff

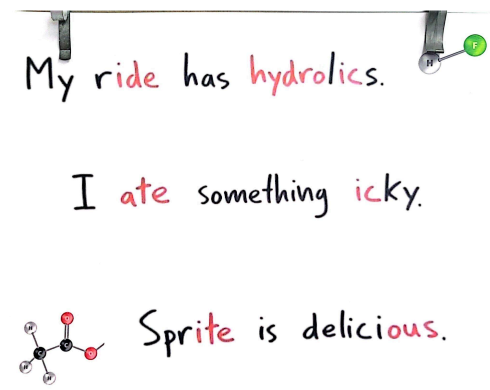


3. Acids/Bases

A. Properties/Uses

Acids	**Base**
Taste sour	Taste bitter
Feels like a burn	Feels slippery
Turns litmus red	Turns litmus blue
Corrosive to metals/stone	Corrosive to organic tissues
Electrolyte	Electrolyte
Neutralize bases	Neutralize acids
pH < 7	pH > 7
Increase H_3O^+ in aqueous solution	Increase OH^- in aqueous solution
Uses: Stomach acid, citrus, carbonic acid, acid rain	Uses: Cleaners, soaps, drain opener, antacids

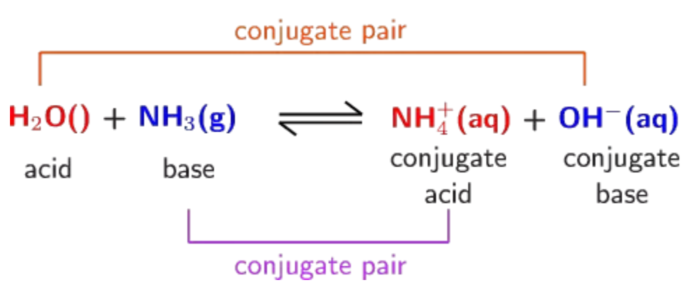
B. Naming Acids



a. Binary: Hydro - (root) -ic acid
   i. Ex: $HF_{(aq)}$ - Hydrofluoric acid
b. Ternary:
   i. -ite: (root) -ous acid
      1. Ex: $H_2SO_{3(aq)}$ - Sulfurous acid
   ii. -ate: (root) -ic acid
      1. Ex: $H_2SO_{4(aq)}$ - Sulfuric acid

C. Definitions

a. Svante Arrhenius
   i. Acids: Donate H^+ (proton)
      1. Ex: HCl, H_2SO_4, H_2CO_3
   ii. Bases: Donate OH^-
      1. Ex: NaOH, KOH, $Ca(OH)_2$
b. Bronsted-Lowry
   i. Acids: Donate H^+
      1. Same examples
   ii. Bases: Take H^+
      1. Ex: NH_3
   iii. Conjugate acid/base pairs
      1. Ex: Pictured at right



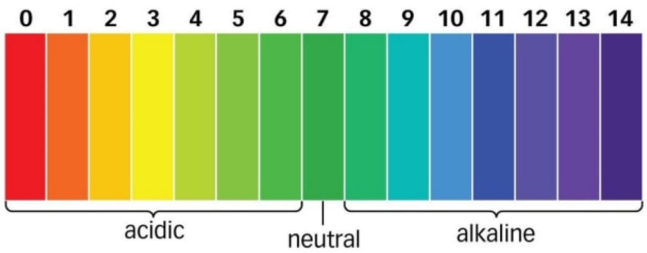
c. Lewis
   i. Acids: Take electron pairs. Electrophile
   ii. Bases: Donate electron pairs. Nucleophile

D. Self ionization of water - Acid/base interactions happen naturally in water

a. $H_2O + H_2O \longleftrightarrow H_3O^+ + OH^-$
b. $K_w = [H_3O^+] \times [OH^-] = [10^{-7}] \times [10^{-7}] = 10^{-14}$

E. pH scale

a. Logarithmic scales - pH is base 10, meaning each step on the scale is 10x more acidic/basic than the previous. Ex: pH 5 is 10x more acidic than pH 6.



b. $pH = -\log[H_3O^+]$

c. $pOH = -\log[OH^-]$

d. $pH + pOH = 14$

F. Strong vs weak acids/bases

a. Strong acids - Assume the following reaction: $HA + H_2O \rightarrow H_3O^+ + A^-$

i. Dissociation constant K_a = [products]/[reactants] = $[H_3O^+][A^-] / [HA][H_2O]$

ii. For strong acids, $K_a >> 1$. Assume that all HA becomes H_3O^+, so $[HA] = [H_3O^+]$. May use [HA] in calculating pH values.

iii. 7 common strong acids: HCl, HBr, HI, HNO_3, $HClO_3$, $HClO_4$ and H_2SO_4

b. Strong bases

i. Exact same assumption as for strong acids. Kb >> 1

ii. Assume that all base dissociates into OH-. [Base] = [OH-]. Use [base] to calculate pOH.

c. Weak acids/bases

i. Dissociation constants Ka & Kb > 1 (quite small).

ii. Cannot assume that all acid/base dissociates, therefore cannot directly calculate pH or pOH. (i.e. [HA] does not equal $[H_3O^+]$)

iii. ICE tables & 5% rule: Stands for Initial concentration, change, equilibrium concentration

	$HCN(aq) \rightleftharpoons$	$H^+(aq)$ +	$CN^-(aq)$
Initial concentration (*M*)	0.15	0	0
Change (*M*)	$-x$	$+x$	$+x$
Equilibrium concentration (*M*)	$0.15 - x$	x	x

K_a for HCN = 3.5×10^{-4} Solve for [H+] (x in table above)

d. Relationship between K_a & K_b for an acid and its conjugate base pair-

i. $K_a \times K_b = K_w$

ii. $pK_a + pK_b = 14$

e. Buffers - Once again assume the following reaction: $HA + H_2O \rightarrow H_3O^+ + A^-$

i. Solutions of weak acids / weak bases with one of their salts contain species of both reactants and products. Some acid does not ionize and remains in solution.

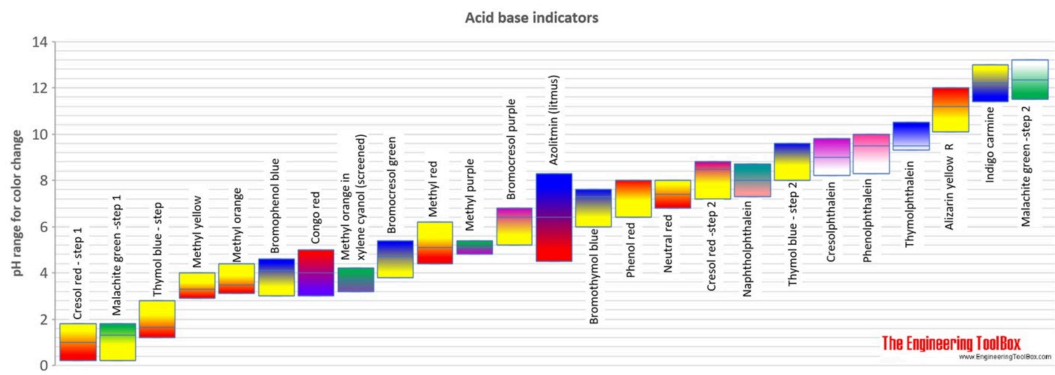
1. Example: A solution of acetic acid ($HC_2H_3O_2$) contains mostly the acid, very little of the conjugate base (acetate ion $C_2H_3O_2^{-1}$). The solution can be “buffered” by adding another source of acetate, like sodium acetate $NaC_2H_3O_2$

ii. Buffered solutions resist changes to changes in pH.

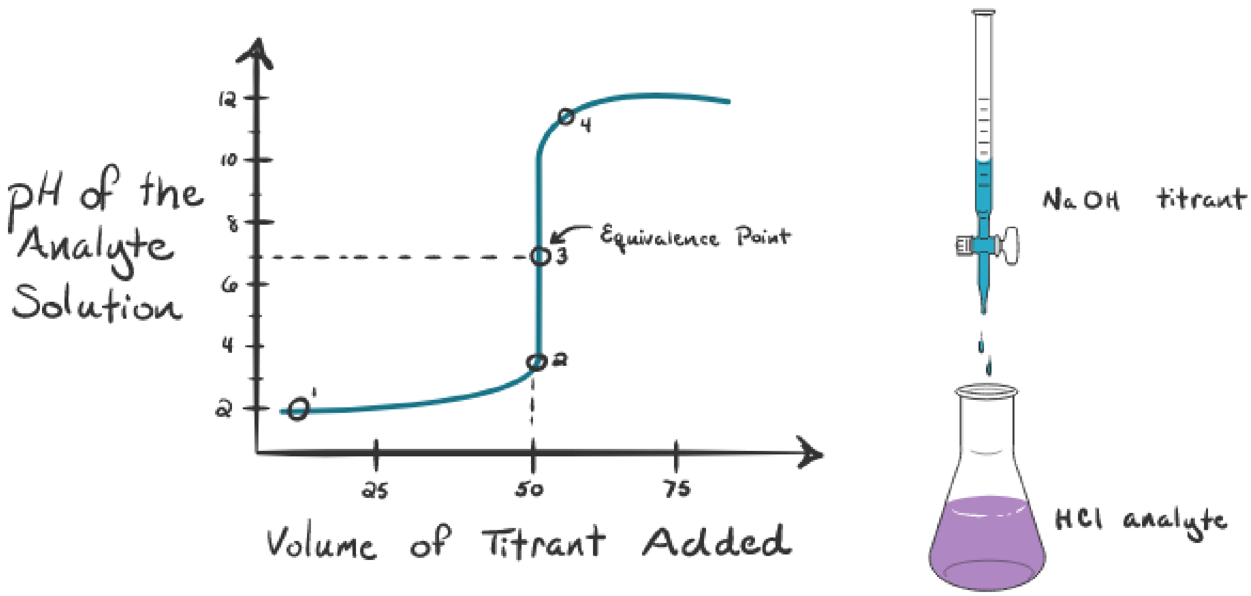
1. Addition of an acid reacts with A-, pushes equilibrium left.
2. Addition of a base reacts with HA

iii. Henderson Hasselbach Equation

f. [pH indicators]() - Usually weak acids/bases. The acid is one color, the conjugate base another. When added to solution, turns color at a specific pH as other acids or bases are added, shifting the equilibrium.



4. Titrations - Technique used to determine the concentration of an unknown solution. Setup pictured below.



A. At the equivalence point, moles of base titrant added = moles of acid analyte below. The equation below applies
   a. MaVa = MbVb
B. Equation above only works when acid and base combine in 1/1 ratio. If not, treat as any other volume solution stoichiometry problem.