Chemistry SAT II Review

Acids and Bases

Properties of acids and bases are caused by ions

1. Hydronium ions (H_3O^+) cause acid properties
2. Hydroxide ions (OH^-) cause base properties

Water - the neutral substance

1. Water is polar
2. The positively charged hydrogens in one water molecule are attracted to the negatively charged oxygens of another
3. Due to this attraction, water ionizes to a very small extent as follows:

$$2H_2O \rightleftharpoons H_3O^+ + OH^-$$

   a. concentration of the ions in pure water - $[H_3O^+]=[OH^-]=10^{-7}M$
   b. importance of ionization of water
      i. since the concentration of hydronium and hydroxide are equal in pure water, water is neutral
      ii. the concentration of ions in pure water is very low because the ions are more likely to combine to form water than water is to ionize
      iii. the limited ionization of water is responsible for the properties of acids and bases

Arrhenius Definition

1. Acids

   a. **Arrhenius Theory** - an acid is a substance that yields hydrogen ions (H^+) as the only positive ions in aqueous solution; the properties of acids are caused by excess hydrogen ions
      i. Acids are polar molecules that contain hydrogen as a metal
      ii. Acids ionize in water to produce hydrogen ions (H^+) or hydronium ions (H_3O^+) [H^+ can't exist alone. It combines with water to form H_3O^+]

$$HCl(g) \rightarrow H^+(aq) + Cl^-(aq) \quad or \quad H_2O(\ell) + HCl(g) \rightarrow H_3O^+(aq) + Cl^-(aq)$$

   b. Other theories explain how substances behave like acids outside of water solution
   c. Examples: [1] HCl; [2] HNO_3; [3] H_2SO_4

2. Base

   a. **Arrhenius' theory** - a base is a substance that yields hydroxide ions as the only negative ions in aqueous solution; the properties of bases are caused by hydroxide ions
      i. Bases are ionic compounds that contain hydroxide as a nonmetal
      ii. Bases dissociate in water to release hydroxide ions

$$NaOH(s) \rightarrow Na^+(aq) + OH^-(aq)$$

   b. Other theories explain how substances behave like bases outside of water solution
   c. Examples: [1] NaOH; [2] NH_4OH; [3] $Ca(OH)_2$ [*NOTE:* Alcohols such as ethanol (C_2H_5OH) are bases because they are not ionic, and do not release OH^- in water]

Properties of Acids and Bases

1. **Acids**
   a. Increase the hydrogen (hydronium) ion concentration of water
   b. Have a pH below 7
   c. Taste sour
   d. Cause color changes in indicators (*indicator - something that reacts with an acid or base to show a definite color change)*
      i. litmus → red
      ii. phenolphthalein → clear
      iii. bromthymol blue → yellow
      iv. methyl orange → red
   e. Conduct electricity
   f. react with active metals to release hydrogen (corrosive)
   g. react with bases to form a salt and water
2. **Bases**
   a. Increase the hydroxide ion concentration of water
   b. Have a pH above 7
   c. Taste bitter
   d. Cause color changes in indicators
      i. litmus → blue
      ii. phenolphthalein → pink
      iii. bromthymol blue → blue
      iv. methyl orange → yellow
   e. Conduct electricity
   f. feel slippery because bases dissolve skin (caustic)
   g. react with acids to form a salt and water

Operational definitions

1. **Arrhenius Theory -**
   a. an acid is a substance that yields hydrogen ions (H^+) as the only positive ions in aqueous solution; the properties of acids are caused by excess hydrogen ions
   b. a base is a substance that yields hydroxide ions as the only negative ions in aqueous solution; the properties of bases are caused by hydroxide ions

2. **Brönsted-Lowry -**
    - a. an acid is any species that can donate a proton to another
        - i. example - when ammonia dissolves in water, water donates a proton to form the ammonium ion; thus water is an acid

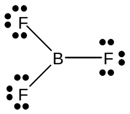
$$NH_3 + H_2O \rightleftarrows NH_4^+ + OH^-$$

    - b. a base is any species (molecule or ion)that can combine with or accept a proton
        - i. example: in the reaction between water and hydrochloric acid, water acts as a Brönsted-Lowry base by accepting a proton

$$HCl + H_2O \rightleftarrows H_3O^+ + Cl^-$$

    - c. **amphoteric** (**amphiprotic**) substances - substances that can act as either Brönsted-Lowry acids or Brönsted-Lowry base depending on what it is reacting with
        - i. examples: H_2O or HSO_4^-
    - d. conjugate acid-base pairs
        - i. in an acid-base reaction, the acid donates a proton to the base
        - ii. the acid, after donating a proton, is capable of accepting its proton back
            - (1) the particle thus formed from the acid is therefor a base
            - (2) the acid and its newly formed base are a conjugate acid-base pair
        - iii. the base, after accepting a proton from an acid, is capable of donating the proton back
            - (1) the particle thus formed from the base is therefor an acid
            - (2) the base and its newly formed acid are a conjugate acid-base pair

3. **The Lewis model**
    - a. Lewis acid = electron pair acceptor
        - i. has an empty atomic orbital that it can use to accept an electron pair from a molecule with a lone pair
        - ii. may be deficient in a pair of electrons
        - iii. examples: H^+ , BF_3



only 6 electrons

    - b. Lewis base = electron pair donor
        - i. molecule with a lone pair

Ionization constants (K_a, K_b, K_w)

K_a

1. The general reaction for an acid dissolved in water is as follows:

$HA(aq)$	+	$H_2O(\ell)$	⇄	$H_3O^+(aq)$	+	$A^-(aq)$
Conjugate Acid 1		Conjugate Base 2		Conjugate Acid 2		Conjugate Base 1

2. The above reaction represents a competition between the bases $H_2O(\ell)$ and $A^-(aq)$ for the proton
   a. if $A^-(aq)$ is a much stronger base than $H_2O(\ell)$, then equilibrium lies to the left and most of the acid will be in the form $HA(aq)$, making $HA(aq)$ a weak acid
   b. If $H_2O(\ell)$ is a much stronger base than $A^-(aq)$, then equilibrium lies to the right and the acid will be largely ionized, making $HA(aq)$ a strong acid
3. The **acid dissociation constant (K_a)** comes from the equilibrium expression for the reaction

$$K_a = \frac{[H_3O^+][A^-]}{[HA]} = \frac{[H^+][A^-]}{[HA]}$$

   a. For acids, the higher the ionization constant (K_a), the stronger the acid
   b. If the acid is ionized completely, [HA] = 0 and K_a is infinite
   c. Ionization constants for very strong acids cannot be calculated; K_a is listed as "very large"
   d. Ionization constants for acids that do *NOT* ionize completely can be calculated as above

K_b

1. The general reaction between a base and water is given by:

$B(aq)$	+	$H_2O(\ell)$	⇌	$BH^+(aq)$	+	$OH^-(aq)$
Base		**Acid**		**Conjugate acid**		**Conjugate base**

2. The equilibrium constant for the general reaction is: $K_b = \frac{[BH^+][OH^-]}{[B]}$,
3. where K_b refers to the reaction of a base with water to form the conjugate acid and the hydroxide ion

K_w - Ionization constant of water (K_w)

1. The ionization equation is $H_2O(\ell) \rightleftarrows H^+(aq) + OH^-(aq)$
2. $K_w = [H^+][OH^-]$
3. at 25°C, $[H^+] = 1 \times 10^{-7}$ moles/L and $[H^+] = [OH^-]$
4. so $K_w = (1 \times 10^{-7}\text{ moles/L})(1 \times 10^{-7}\text{ moles/L}) = 1 \times 10^{-14}\text{ moles}^2/\text{L}^2$
5. The ionization constant is

$$\therefore\ K_w = [H_3O^+][OH^-] = 10^{-14} mol^2 / L^2 \text{ at } 25°C$$

6. significance - in any aqueous solution, no matter what else it contains, at 25°C, the product of $[OH^-]$ and $[H^+]$ is always 1.0×10^{-14}, resulting in three possible situations
   a. neutral solution: $[H^+] = [OH^-]$
   b. acidic solution: $[H^+] > [OH^-]$; and
   c. basic solution: $[H^+] < [OH^-]$
7. Calculating the concentration of hydronium or hydroxide when one of the concentrations are known
   a. If $K_w = [H^+][OH^-] = 1 \times 10^{-14}$ moles/L, then
   b. $[H^+] = \frac{1\times10^{-14}}{[OH^-]}$, and $[OH^-] = \frac{1\times10^{-14}}{[H^+]}$

Acid-base indicators

1. Definition
   a. Acid–Base indicators - compounds whose colors are sensitive to pH
2. Nature
   a. complex molecules that are weak acids (HIn)
   b. they exhibit one color when the proton is attached and a different color when the proton is absent
   c. example: phenolphthalein
      i. colorless as HIn
      ii. pink as In^-
3. Functionning
   a. Equilibrium for hypothetical indicator HIn

$$\text{HIn}(aq) \rightleftharpoons \text{H}^+(aq) + \text{In}^-(aq)$$

   b. In basic solutions, OH^- removes H^+ shifting equilibrium to the right
4. Transition range - pH range over which an indicator changes color

Titration Type	Indicator	Range	Color
			Acid → Base
Strong Acid/Strong Base	methyl red	4.4-6.2	Red → Yellow
	bromthymol blue	6.2-7.6	Yellow → Blue
Strong Acid/Weak Base	methyl orange	3.1-4.4	Red → Yellow
	bromphenol blue	3.0-4.6	Yellow → Red
Weak Acid/Strong Base	phenolphthalein	8.0-10.0	Colorless → Red
	phenol red	6.4-8.0	Yellow → Red

pH

1. Definition - measure of hydronium ion concentration
   ($pH = -\log[H_3O^+]$)
2. Calculating pH

$[H_3O^+]$	$[OH^-]$	pH
10^{-1} M	10^{-13} M	1
10^{-2} M	10^{-12} M	2
10^{-3} M	10^{-11} M	3
10^{-4} M	10^{-10} M	4
10^{-5} M	10^{-9} M	5
10^{-6} M	10^{-8} M	6
10^{-7} M	10^{-7} M	7
10^{-8} M	10^{-6} M	8
10^{-9} M	10^{-5} M	9
10^{-10} M	10^{-4} M	10
10^{-11} M	10^{-3} M	11
10^{-12} M	10^{-2} M	12
10^{-13} M	10^{-1} M	13

Neutralization

1. Definition - reaction between an acid and a base to produce a salt and water
2. Example

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O$$

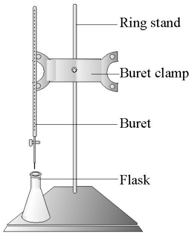
3. Ions present during neutralization reaction above
   a. $HCl + H_2O \rightarrow H_3O^+ + Cl^-$
   b. $NaOH + H_2O \rightarrow Na^+ + OH^- + H_2O$
   c. $NaCl + H_2O \rightarrow Na^+ + Cl^- + H_2O$

$$H_3O^+ + \underbrace{Cl^- + Na^+}_{\text{spectator ions}} + OH^- \rightarrow \underbrace{Na^+ + Cl^-}_{\text{spectator ions}} + 2H_2O$$

NET REACTION: $H_3O^+ + OH^- \rightarrow 2H_2O$

Titration

1. Definition - method of determining the concentration of an acid or base by neutralizing it with a solution of known concentration
2. Procedure



   a. Place a measured amount of acid or base of unknown concentration in a flask and add two drops of phenolphthalein
   b. Use a ring stand with a buret clamp and a buret as shown in the diagram to the right. Fill the buret with a standard solution (an acid or base of known concentration)
      i. the buret is used to dispense the standard solution and measure the amount dispensed
   c. Hold the flask containing the acid or base of unknown concentration under the buret. Run the standard solution slowly into the flask, mixing occasionally by swirling. When the color begins to change on contact with the standard solution, add the standard solution one drop at a time until one final drop causes a complete and permanent color change.
   d. Determine the volume of standard solution used
   e. Calculate the concentration of the unknown solution using the data you gathered and the equation below
3. Calculation

$$\mathbf{M_a \times V_a = M_b \times V_b}$$

4. Diprotic and triprotic acids/dihydroxy and trihydroxy bases
   a. During a neutralization reaction each hydrogen ion (hydronium ion) is neutralized by one hydroxide ion
   b. Therefore, during a titration, the concentration of hydrogen ions and hydroxide ions is more important than the concentration of the acid or base, so it is necessary to determine the effective concentration due to these ions
   c. Effective concentration
      i. Polyprotic acids
         (1) Sulfuric acid (H_2SO_4) is diprotic: It forms 2 *mol* of hydrogen ions (protons) per *mol* of acid $H_2SO_4(aq) \rightarrow 2H^+(aq) + SO_4^{2-}(aq)$
         (2) The effective concentration of 0.2M H_2SO_4 is 0.4M in titration problems

$$\underline{\underline{M_{AE} = M_A \times n_H}}$$

M_{AE} = effective concentration of acid
M_A = concentration of acid
n_H = number of hydrogens

$$\mathbf{M_A} = \frac{\mathbf{M_{AE}}}{n_\mathbf{H}}$$

      ii. Polyhydroxy bases
         (1) Calcium hydroxide [$Ca(OH)_2$] is dihydroxy: It forms 2 *mol* of hydroxide ions per *mol* of base $Ca(OH)_2(aq) \rightarrow Ca^{2+}(aq) + 2OH^-(aq)$
         (2) The effective concentration of 0.25M H_2SO_4 is 0.5M in titration problems

$$\underline{\underline{M_{BE} = M_B \times n_{OH}}}$$

M_{BE} = effective concentration of base
M_B = concentration of base
n_{OH} = number of hydroxides

$$\mathbf{M_B} = \frac{\mathbf{M_{BE}}}{n_\mathbf{OH}}$$

Naming Acids

1. Binary acids - prefix HYDRO and suffix IC
    a. HCl = Hydrochloric
    b. HBr = Hydrobromic
2. Oxyacids (containing polyatomic ions with oxygen)
    a. Rules
        i. most common number of oxygens = suffix IC
            $HClO_3$ = chloric acid
        ii. one more than most common = prefix PER and suffix IC
            $HClO_4$ = perchloric acid
        iii. one less than most common = suffix OUS
            $HClO_2$ = chlorous acid
        iv. two less than most common = prefix HYPO and suffix OUS
            HClO = hypochlorous acid
    b. Easiest to do by using relationship between acid name and polyatomic ion name

oxidation state	polyatomic ion			acid name	
	example	prefix	suffix	prefix	suffix
two less than most common	ClO^{-1}	hypo	ite	hypo	ous
one less than most common	ClO_2^{-1}	-	ite	-	ous
most common	ClO_3^{-1}	-	ate	-	ic
one more than most common	ClO_4^{-1}	hyper	ate	per	ic

Strength of Acids and Bases

1. Acids - the stronger the nonmetal part of the acid is, the more polar the acid is and the more acidic it is; (Table L lists the relative strength of acids; stronger acids have larger K_as)
   a. Strong acids
      i. All acids of Group VIIA nonmetals: [1] HCl; [2] HI
      ii. Industrial acids: [1] H_2SO_4; [2] HNO_3
   b. Weak acids: [1] Organic acids (containing carbon); [2] H_3PO_4
2. Bases - the stronger the metal part of the base is, the more ionic the base is and the more basic
   a. Strong bases - Bases of Group IA and IIA metals: [1] NaOH; [2] $Ca(OH)_2$
   b. Weak bases - all other bases such as [1] NH_4OH; or [2] $Al(OH)_3$

Hydrolysis

1. Definition - reverse of neutralization (a salt reacts with water to produce an acid and a base)
2. Example - $NH_4CH_3COO + H_2O \rightarrow NH_4OH + HCH_3COO$
3. Explanation
   a. there are a small number of hydronium and hydroxide ions in pure water
   b. the metal ions from the salt can combine with hydroxide ions to form a base
      i. if the base formed is strong, it dissociates back into ions
      ii. if the base formed is weak, it does not dissociate
   c. the nonmetal ions from the salt can combine with hydronium ions to form an acid and water
      i. if the acid formed is strong, it ionizes again
      ii. if the acid formed is weak, it does not ionize
4. Significance - salts may not be neutral
   a. A salt of a strong acid and a weak base is ACID
      i. the salt dissolves in water to form a strong acid which reionizes releasing hydroniums
      ii. the salt dissolves in water to produce a weak base which does not dissociate so hydroxides are removed from solution
   b. A salt of a weak acid and strong base is a BASE
      i. the salt dissolves in water to form a weak acid which does not ionize so it removes hydroniums from solution
      ii. the salt dissolves in water to form a strong base which dissociates releasing hydroxide

Acid and base anhydrides

1. Definition - Acid or base minus water
2. Acid Anhydride
   a. Examples

$$\begin{array}{lll} H_2CO_3 & H_2SO_3 & 2HNO_3 \\ \underline{-\ H_2O\ } & \underline{-\ H_2O\ } & \underline{-\ H_2O\ } \\ CO_2 & SO_2 & N_2O_3 \end{array}$$

   b. Definition - nonmetallic oxides
   c. Acid anhydrides as a source of acid rain

Environmental Source	Nonmetallic Oxide	Acid Formed
Car exhaust	Carbon dioxide Nitrogen oxides	Carbonic acid Nitric acid
Coal	Sulfur dioxide Carbon dioxide	Sulfurous acid Carbonic acid
Smelters	Sulfur dioxide Sulfur trioxide	Sulfurous acid Sulfuric Acid
Volcanoes	Sulfur dioxide	Sulfurous acid
Lightning	Nitrogen oxides	Nitric acid

3. Base Anhydride
   a. Examples

$$\begin{array}{lll} 2NaOH & Ca(OH)_2 & 2Al(OH)_3 \\ \underline{-\ H_2O\ } & \underline{-\ H_2O\ } & \underline{-\ 3H_2O\ } \\ Na_2O & CaO & Al_2O_3 \end{array}$$

   b. Definition - metallic oxides
   c. Base anhydrides in the environment: oxides such as lime (CaO) are used to neutralize acid soil