Chapter 19- Acids, Bases and Salts.   GA chemistry standards:
SC7.b: compare contrast and evaluate the nature of acids and bases
SC7.b.1: Arrhenius, Bronsted-Lowery acid/bases
SC7.b.2: strong vs. weak acids/bases in terms of percent dissociation
SC7.b.3: hydronium ion concentration (hydrogen ion)
SC7.b.4: Acid-Base neutralization

19.1 Acid-Base Theories

- Properties of Acids
    - Sour taste
    - React with metals to produce hydrogen gas
    - Aqueous solutions conduct electricity and are call electrolytes
    - Will change the color of acid base indicator
    - Low pH (below 7)
- Properties of Bases
    - Bitter taste
    - Feel slippery
    - Aqueous solutions conduct electricity and are call electrolytes
    - Will change the color of acid base indicator
    - High pH (above 7)
- Arrhenius Acids are defined as hydrogen-containing compounds that ionize to yield hydrogen ions (H^{+1}) in aqueous solutions
    - Monoprotic acid are acids that contain ONE ionizable hydrogen. Examples: HCl or HNO_3
    - Diprotic acid are acids that contain TWO ionizable hydrogen. Examples: H_2S or H_2SO_4
    - Triprotic acids are acids that contain THREE ionizable hydrogen. Examples: H_3PO_4
- Arrhenius bases are defined as compounds that ionize to yield hydroxide ions (OH^{1-}) in aqueous solutions
- Bronsted-Lowery defines an acid as a Hydrogen-Ion donor
- Bronsted-Lowery defines a base as a hydrogen-ion acceptor
    - Conjugate acid is the particle formed when a base gains a hydrogen ion
    - Conjugate base is the particle formed when an acid looses a hydrogen ion
    - Conjugate acid-base pair consist of two substances that are related by the loss or gain of a single hydrogen ion

$$\underset{\text{Base}}{NH_3(aq)} + \underset{\text{Acid}}{H_2O(l)} \rightleftharpoons \underset{\text{Conjugate acid}}{NH_4^+(aq)} + \underset{\text{Conjugate base}}{OH^-(aq)}$$

    - Conjugate acid-base pairs: NH_3 & NH_4^{+1}, H_2O & OH^{-1}
    - **If the compounds differ by more that one hydrogen ion (or any other element) they CAN NOT be classified as conjugate acid-base pairs.**
- Lewis acid is a substance that can accept a pair of electrons to form a covalent bond
- Lewis base is a substance that can donate a pair of electrons to form a covalent bond

Type	Acid	Base
Arrhenius	H^+ producer (donor)	OH^{-1} producer
Bronsted-Lowery	H^+ producer (donor)	H^+ acceptor
Lewis	Electron-pair acceptor	Electon-pair donor

19.2 Hydrogen Ions and Acidity

- Self-ionization is where water molecules production ions

$$H_2O(l) \rightleftharpoons \underset{\text{Hydrogen ion}}{H^+(aq)} + \underset{\text{Hydroxide ion}}{OH^-(aq)}$$

- Neutral solution is where the [H^{+1}] and [OH^{-1}] are equal ([]=concentration of)

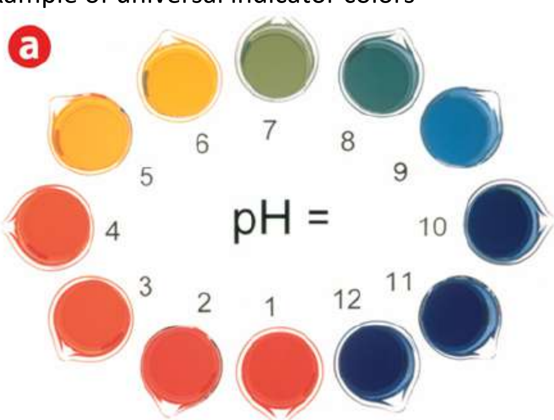
- **For aqueous solutions, the product of the hydrogen-ion concentration and the hydroxide-ion concentrations equals 1.0 X 10^{-14}.**
- Ion-product constant for water (K_w) is the product of the concentrations of the hydrogen ions and hydroxide ions in water.
  - $K_w = [OH^{-1}]\ [H^{+1}] = 10^{-14}$
- Acidic solution is one in which $[H^{+1}]$ is greater than $[OH^{-1}]$, the $[H^{+1}]$ is greater than 1.0 X 10^{-7}.
- Basic solution is one in which $[H^{+1}]$ is less than $[OH^{-1}]$, the $[H^{+1}]$ is less than 1.0 X 10^{-7}.
- pH of a solution is the negative logarithm of the hydrogen-ion concentration
- $pH = -\log[H^{+1}]$
- **A solution in which $[H^{+1}]$ is greater than 1.0 X 10^{-7}M has a pH less than 7.0 and is acidic**
- **A solution in which $[OH^{-1}]$ is less than 1.0 X 10^{-7}M has a pH greater than 7.0 and is basic**
- **Recall that M stands for molarity and is moles/liter and is used to represent concentration**
- pOH of a solution is the negative logarithm of the hydroxide-ion concentration
- $pH = -\log[OH^{-1}]$

Table 19.5

Relationship among $[H^+]$, $[OH^-]$, and pH

	$[H^+]$ (mol/L)	$[OH^-]$ (mol/L)	pH	Aqueous system
	1×10^{0}	1×10^{-14}	0.0	1*M* HCl
	1×10^{-1}	1×10^{-13}	1.0	0.1*M* HCl
	1×10^{-2}	1×10^{-12}	2.0	Gastric juice, Lemon juice
Increasing acidity	1×10^{-3}	1×10^{-11}	3.0	
	1×10^{-4}	1×10^{-10}	4.0	Tomato juice
	1×10^{-5}	1×10^{-9}	5.0	Black coffee
	1×10^{-6}	1×10^{-8}	6.0	Milk
Neutral	1×10^{-7}	1×10^{-7}	7.0	Pure water
	1×10^{-8}	1×10^{-6}	8.0	Blood
	1×10^{-9}	1×10^{-5}	9.0	Sodium bicarbonate, sea water
	1×10^{-10}	1×10^{-4}	10.0	
Increasing basicity	1×10^{-11}	1×10^{-3}	11.0	Milk of magnesia, Household ammonia
	1×10^{-12}	1×10^{-2}	12.0	Washing soda
	1×10^{-13}	1×10^{-1}	13.0	0.1*M* NaOH
	1×10^{-14}	1×10^{0}	14.0	1*M* NaOH

- **An indicator is a valuable tool for measuring pH because its acid form and base form have different colors in solution**
- Example of universal indicator colors



Using pH or pOH to calculate concentration

- **If pH or pOH is known they can be used to calculate the concentration of hydrogen ions or hydroxide ions in solution.**
  - $[H^{+1}] = 10^{-pH}$

- $[OH^{-1}] = 10^{-pOH}$
- pH + pOH = 14
- $[OH^{-1}]\ [H^{+1}] = 10^{-14}$

- Example find the pH, pOH and the $[OH^{-1}]$ if the $[H^{+1}] = 1.23 \times 10^{-2}$
  - $[H^{+1}] = 1.23 \times 10^{-2}$
  - pH = -log $(1.23 \times 10^{-2}) = 1.91$
  - pOH = 14 – 1.91 =12.09
  - $[OH^{-1}] = 10^{(-12.09)} = 8.13 \times 10^{-13}$

Acid	Neutral	Base
pH less than 7 $[H^{+1}]$ is greater than $1.0 \text{ X } 10^{-7}$M	pH = 7 $[H^{+1}]$ is equal to $1.0 \text{ X } 10^{-7}$M	pH greater than 7 $[H^{+1}]$ is less than $1.0 \text{ X } 10^{-7}$M

19.3 Strengths of Acids and Bases

- Strong acids are complete ionized in aqueous solutions
  - Example: $HCl \rightarrow H^{+1} + C^{-1}$ forms 100% of the expected ions
  - For strong acid the molarity equals the concentration of hydrogen ions
- Weak acids ionizes only slightly in aqueous solutions
  - Example: $HC_2H_3O_2 \rightarrow H^{+1} + C_2H_3O_2^{-1}$ forms 10-15% of the expected ions
- Strong bases are complete ionized in aqueous solutions
  - Example: $NaOH \rightarrow Na^{+1} + OH^{-1}$ forms 100% of the expected ions
  - For strong bases the molarity equals the concentration of hydroxide ions
- Weak bases ionizes only slightly in aqueous solutions
  - Example: $NH_3 + H_2O \rightarrow NH_4^{+1} + OH^{-1}$ forms 10-15% of the expected ions

19.4 Neutralization reactions

- Neutralization reaction is a reaction in which an acid and a base react in an aqueous solution to produce a salt and water.
- **In general the reaction of an acid with a base produces water and one of a class of compounds called salts.**
- Equivalence point is when the number of moles of hydrogen ions equals the number of moles of hydroxide ions
- Titration is the process of adding a known amount of solution of known concentration to determine the concentration of another solution.
- Standard solution is the solution of known concentration
- Titration continues until the indicator shows that neutralization had just occurred.
- End point is the point at which the indicator changes color
- **The point of neutralization is the end point of the titration.**

Titration calculations

1. Start by writing a BALANCED chemical reaction if one is not given
2. Find the number of moles of standard solution used (Molarity x Liters = moles)
3. Use stoichiometry to see how many moles of the other reactant were neutralized
4. Use the known volume of second reactant to calculate the concentration (Molarity = moles/liters). Or use the concentration to calculate the volume needed.

Example1: How much of 0.5M HNO_3 in mL is necessary to titrate 25.0 mL of a 0.05 M $Ca(OH)_2$?

1. Balanced equation: $2\ HNO_3 + Ca(OH)_2 \rightarrow 2\ H_2O + Ca(NO_3)_2$
2. Moles of standard solution used:
   a. $25.0\ mL \frac{1\ L}{1000\ mL} = 0.0250\ L$ **b.** $0.05\ M\ Ca(OH)_2 = 0.05 \frac{mol}{L}\ Ca(OH)_2$
   c. $0.05 \frac{mol}{L}\ Ca(OH)_2 \times 0.0250\ L = 0.00125\ mol\ Ca(OH)_2$
3. Stoichiometry: $0.00125\ mol\ Ca(OH)_2\ \frac{2\ mol\ HNO_3}{1\ mol\ Ca(OH)_2} = 0.0025\ mol\ HNO_3$
4. Volume: $\frac{0.0025\ mol}{0.5\ \frac{mol}{L}} HNO_3 = 0.005 L HNO_3 \frac{1000 mL}{1L} = 5.0\ mL\ HNO_3$

Example2: What is the concentration of 22.0 mL HNO_3 that is titrated 25.0 mL of a 0.15 M $Ca(OH)_2$?

1. Balanced equation: $2\ HNO_3 + Ca(OH)_2 \rightarrow 2\ H_2O + Ca(NO_3)_2$
2. Moles of standard solution used:
   a. $25.0\ mL\ \frac{1\ L}{1000\ mL} = 0.0250\ L$ **b.** $0.15\ M\ Ca(OH)_2 = 0.15\frac{mol}{L}\ Ca(OH)_2$
   c. $0.15\frac{mol}{L}\ Ca(OH)_2 \times 0.0250\ L = 0.00375\ mol\ Ca(OH)_2$
3. Stoichiometry: $0.00375\ mol\ Ca(OH)_2\ \frac{2\ mol\ HNO_3}{1\ mol\ Ca(OH)_2} = 0.0075\ mol\ HNO_3$
4. concentration: $22.0\ mL\ \frac{1\ L}{1000\ mL} = 0.0220\ L$
   $\frac{0.0075\ mol}{0.0220\ L} HNO_3 = 0.341\ MHNO_3$

19.5 Salts in solution

- Salt hydrolysis is where the cations or anions of a dissociated salt remove hydrogen ions from (or donate hydrogen ions to) the water.
- **In general, salts that produce acidic solution contain positive ions that release protons to water. Salts that produce basic solutions contain negative ions that attract protons from water.**
- Buffer is a solution in which the pH remains relatively constant when small amounts of acid or bases are added.
- **A buffer is a solution of weak acid and one of its salts, or a solution of a weak base and one of its salts.**
- A buffer solute is better able to resist drastic changes in pH than is pure water
- Buffer capacity is the amount of acid or base that can be added to a buffer solution before a significant change in pH occurs.
- Two buffer systems are crucial in maintain human blood pH.

EQUATION THAT YOU MUST KNOW THE DAY OF THE TEST

pH= -log $[H^{+1}]$ pOH= -log $[OH^{-1}]$ $[H^{+1}]=10^{-pH}$ $[OH^{-1}]= 10^{-pOH}$ pH + pOH = 14 $[OH^{-1}]\ [H^{+1}]=10^{-14}$	$Molarity = \frac{mol}{L}$ $1000\ mL = 1\ L$