Ms. Randall
Regents Chemistry
Unit 10: Acids, Bases, and Salts

Unit Objectives:

- Identify substances as Arrhenius acids or Arrhenius bases
- Identify solutions as acid, base, or neutral based upon the pH
- Interpret changes in acid-base indicator color
- Write simple neutralization reactions when given the reactants
- Calculate the concentration or volume of a solution, using titration data
- Define an electrolyte
- Determine the acidity or alkalinity of an aqueous solution can be measured by its pH value.
- Relate the pH scale to hydronium ion concentration.
- Explain behavior of many acids and bases by the Arrhenius theory.
- Define Arrhenius acids and bases
- Explain the process of neutralization
- Describe alternate acid-base theories.
- Perform a simple Titration process in which a volume of a solution of known concentration is used to determine the concentration of another solution.

Focus Questions for the Unit:

- How can precise concentrations be made?
- How are electrolytes described, in terms of formula, properties, concentration, and pH?

YOU SHOULD BE ABLE TO ANSWER THESE *IN DETAIL* BY THE END OF THE UNIT

Define the following vocabulary:

Amphoteric
Arrhenius acid
Arrhenius base
Bronsted-Lowry acid
Bronsted-Lowry base
Electrolyte
hydronium ion
hydroxide ion
indicator (acid/base)
neutralization
pH scale
titration

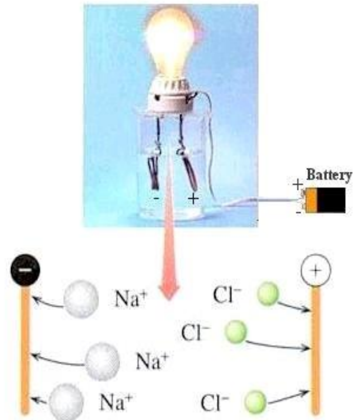
Lesson 1: Chapter Diary 15

Objective: To summarize concepts related to acids and bases

Directions: After reading the Chapter diary 15 answer the questions in your workbook.

<u>Lesson 2: Acids and Bases are Electrolytes</u>

<u>Objective:</u> To compare and contrast the three types of electrolytes: acids, bases and salts.



For ionic compounds that dissolve in water, describing them as "**electrolytes**" is appropriate, since the crystal will dissociate (fall apart in water). **An electrolyte is a substance that dissolves in water and forms a solution capable of conducting an electric current.** The ability of a solution to conduct an electric current depends upon the concentration of ions that are present. If a substance is referred to as an "**electrolyte**," then it must be able to dissolve in water. When it does dissolve, it forms ions in solution. As a result it can conduct electricity. "-Lyte" is Latin for "break" and "electro-"is Latin for "electricity." Putting that all together then, an "electrolyte" is a substance that "breaks" into ions when dissolved in water, and as a result, conducts "electricity."

3 Types of Electrolytes:

Electrolytes are classified according to the type of ions formed by the substance when it dissolves. Svante Arrhenius was a Swedish chemist who first studied this branch of chemistry, and developed the following categories:

1. <u>Arrhenius **Acid**</u> – a substance that dissolves to form H^{+1} ion as the ONLY positive ion in solution.

2. <u>Arrhenius **Base**</u> - a substance that dissolves to form $(OH)^{-1}$ ion as the ONLY negative ion in solution.

3. <u>**Salts**</u> – a substance that dissolves to form a positive ion other than H^{+1} and a negative ion other than $(OH)^{-1}$. Salts are ionic substances built out of a metal cation and a nonmetal anion.

<u>Dissolving in water</u>	<u>Type of electrolyte</u>	<u>Why?</u>
1. $HCl_{(s)} \rightarrow H^{+1}_{(aq)} + Cl^{-1}_{(aq)}$	acid	H^{+1} is the only positive ion in solution
2. $NaOH_{(s)} \rightarrow Na^{+1}_{(aq)} + (OH)^{-1}_{(aq)}$	base	$(OH)^{-1}$ is the only negative ion in solution
3. $K(NO_3)_{(s)} \rightarrow K^{+1}_{(aq)} + (NO_3)^{-1}_{(aq)}$	salt	Positive and negative ions other than H^{+1} and $(OH)^{-1}$ are present
4. $NH_{3(aq)} \rightarrow NH_4^{+1}{}_{(aq)} + (OH)^{-1}_{(aq)}$	base	$(OH)^{-1}$ is the only negative ion in solution

Lesson 3: Characteristics of Acids, Bases and salts

Objective: To compare and contrast the chemical and physical properties of acids and bases

Acids dissociate in water to form **H^{+1} ions** (sometimes referred to as H_3O^{+1} or "hydronium" ion).

Example: HCl is a **monoprotic acid** because when it is dissolved in water **1 mole of H^{+1}** ions are present.

- **Monoprotic**: **mono** means 1, **protic** for "proton" or H+1 ion

There are also **diprotic, triprotic, etc.** acids depending on how many moles of H^{+1} ions are present in solution

Examples:

$HCl \rightarrow \mathbf{H^{+1}} + Cl^{-1}$ monoprotic acid = 1 H^+ ion produced

$H_2(SO_4) \rightarrow \mathbf{2\ H^{+1}} + (SO_4)^{2-}$ diprotic acid = 2 H^+ ions produced

Properties

- Dilute solutions of acids have a **SOUR** taste (examples: citric acid in OJ & lemons/limes, acetic acid in vinegar, carbonic acid in soda, boric acid used as eye-washing solution).
- Concentrated solutions of acids **BURN** skin & **EAT HOLES** in clothing
- Aqueous solutions of acids are **ELECTROLYTES** (substances that conduct electric current); the more **H^+ IONS** in solution, the more **CONDUCTIVE** the acid (weak acids = poor conductors, strong acids = good conductors)
- Acids **REACT WITH BASES TO FORM NEUTRAL** solutions (Acid + Base → Salt + Water); involve **DOUBLE REPLACEMENT** reactions
- Acids **REACT WITH CERTAIN METALS TO PRODUCE HYDROGEN GAS (H_2)** in a **SINGLE REPLACEMENT** reaction (Table J) any **METAL LOCATED ABOVE H_2 WILL REACT WITH AN ACID** to produce H_2 gas and a salt
- Acids cause acid-base **INDICATORS** to **CHANGE COLOR** (Ex: **LITMUS** turns **RED** & **PHENOLPHTHALEIN** is **COLORLESS** in presence of acid)
- Acids have **pH VALUES < 7** (fall on **LOWER** end of **pH SCALE**)
- **GENERAL FORMULA = HA** or **HX** (Where X = neg. ion such as Cl^-)

Table K
Common Acids

Formula	Name
$HCl(aq)$	hydrochloric acid
$HNO_3(aq)$	nitric acid
$H_2SO_4(aq)$	sulfuric acid
$H_3PO_4(aq)$	phosphoric acid
$H_2CO_3(aq)$ or $CO_2(aq)$	carbonic acid
$CH_3COOH(aq)$ or $HC_2H_3O_2(aq)$	ethanoic acid (acetic acid)

Bases dissociate in water to form the hydroxide ion, **$(OH)^{-1}$**.

Base Example: NaOH is a **monohydoxy base** because when it is dissolved in water **1 mole of (OH)-1** ions are present in solution.
Monohydroxy: **mono** means 1, **hydroxy** for "hydroxide ion"
There are also **dihydroxy, trihydroxy, etc.** bases depending on how many moles of $(OH)^{-1}$ ions are present in solution

Examples:

$NaOH \rightarrow Na^{+1} + \mathbf{OH^{-1}}$	monohydroxy base = 1 OH- ion produced
$Ca(OH)_2 \rightarrow Ca^{2+} + \mathbf{2\ OH^{-1}}$	dihydroxy base = 2 OH- ions produced

Properties

- Dilute solutions of bases have a **BITTER** taste
  (Examples: antacid for upset stomach, soap, *drano—don't taste this one)
- Bases have a **SLIPPERY** or **SOAPY** feel
- Aqueous solutions of bases are **ELECTROLYTES**
  (substances that conduct electric current) – the more **OH^- IONS** in solution, the more **CONDUCTIVE** the base (weak bases are poor conductors, strong bases are good conductors)
- Bases **REACT WITH ACIDS TO FORM NEUTRAL** solutions
  (Acid + Base → Salt + Water) – Neutralization reactions are a specific type of **DOUBLE REPLACEMENT** reaction
- Bases cause acid-base indicators to change color (Example: **LITMUS TURNS BLUE** & **PHENOLPHTHALEIN IS PINK** in the presence of a base)
- Bases have **pH VALUES > 7** (fall on **HIGHER** end of **pH SCALE**)
- **GENERAL FORMULA = XOH** (Where X = positive ion such as Na^+)

****NONEXAMPLES OF BASES:**
There are some compounds that do contain –OH that ARE NOT bases.
If covalent bonds are involved, they cannot ionize and are incapable of producing ions. For example, **ALCOHOLS** (which have covalently bonded long carbon chains) **ARE NOT BASES!**

Examples: CH_3CH_2OH (ethanol) – covalent due to carbon chain →
Don't produce ions → nonelectrolytes! Bases **must have ionic bonds**!

Table L
Common Bases

Formula	Name
$NaOH(aq)$	sodium hydroxide
$KOH(aq)$	potassium hydroxide
$Ca(OH)_2(aq)$	calcium hydroxide
$NH_3(aq)$	aqueous ammonia

Amphoteric(amphiprotic) substances

- Have characteristics of both an acid and a base and capable of reacting as either
- Water is amphoteric

$$H_2O \quad + \quad H_2O \rightarrow H_3O^+ \quad + \quad OH^-$$

- Amphoteric substances must have at least one hydrogen in their formula

Example: Which of the following substances is amphoteric?

a.) H_2SO_3

b.) KCl

c.) HSO_3^-

d.) Br^-

Answer: a

Properties of Salts:

- Defined as **NEUTRAL IONIC SUBSTANCES** that have **POSITIVE IONS OTHER THAN HYDROGEN** and **NEGATIVE IONS OTHER THAN HYDROXIDE**; composed of a positively charged metal or polyatomic ion AND a negatively charged nonmetal or polyatomic ion (with the exception of $-OH^-$ which would then make it a base)

- Examples of salts → **LiBr, KI, $CaCl_2$, $NaNO_3$**

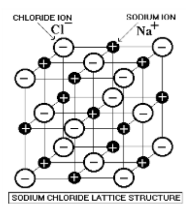
- Salts are **FORMED FROM NEUTRALIZATION REACTIONS** and are **NEUTRAL**

acid + base → **SALT** + water

- Exception to the rule: **NH_4Cl** is the salt of a weak acid

$$NH_4Cl \rightarrow NH_4^+ + Cl^-$$

NH_4^+ is a weak acid, which means it produces H^+ in solution

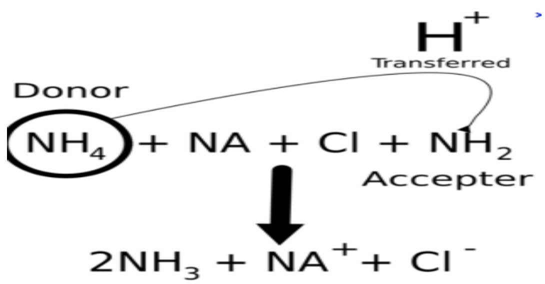


ACIDS, BASES, & SALTS ARE ALL ELECTROLYTES (in SOLUTION)

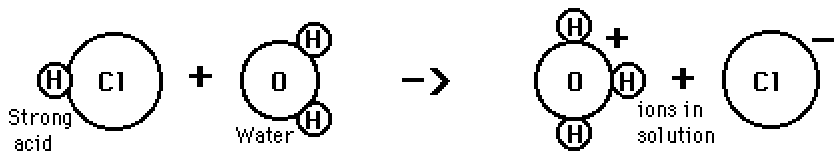
Lesson 4: Bronsted-Lowry and Buffers

Objective: To relate weak acids and bases to buffering systems (Le Chatelier's revisited!)

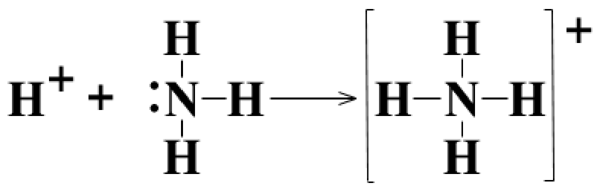
This alternate theory explains the behavior of weak acids and bases:



BRONSTED-LOWRY ACIDS are the **PROTONS DONORS** (or H+ DONORS).
They prefer to GIVE AWAY A PROTON (H^+) to the base.

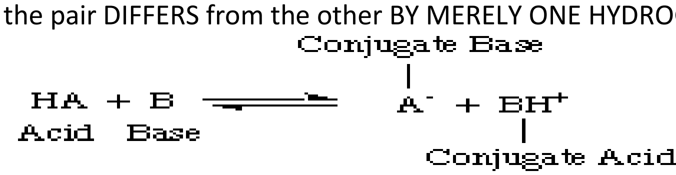


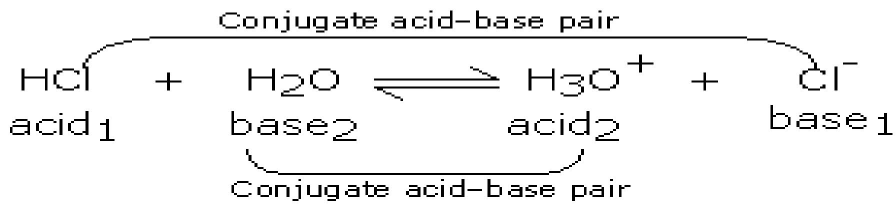
BRONSTED-LOWRY BASES are the **PROTON ACCEPTORS** (or H+ ACCEPTORS). They prefer to ACCEPT/RECEIVE A PROTON (H^+) from the acid.



In the following reaction, let's label the acids and the bases according to Bronsted-Lowry:

We will use BRACKETS to connect one side of the reaction to the other to represent the ACID-BASE CONJUGATE PAIRS (each member within the pair DIFFERS from the other BY MERELY ONE HYDROGEN)



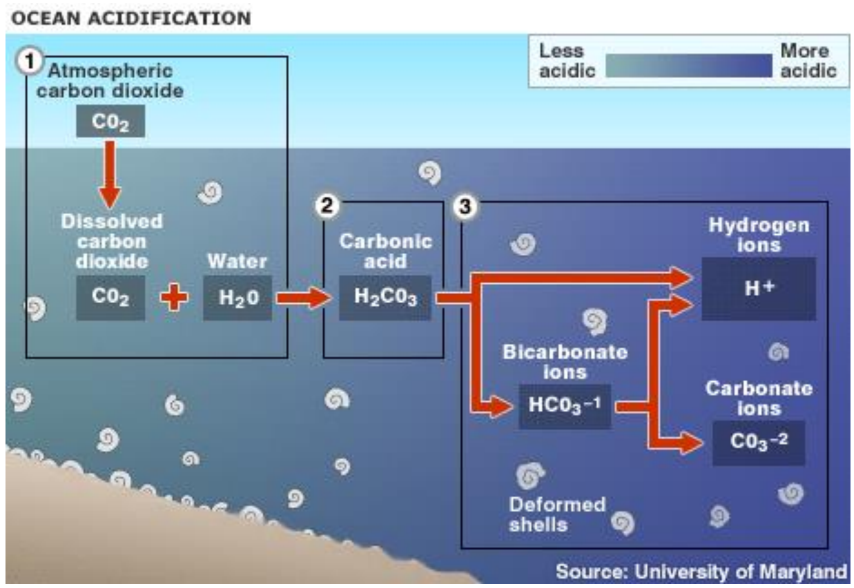


Notice: The acid always has one more H+ than the base!

Summary:

- Explains behavior of weak acids and bases
- Acids hydrogen donors
- Bases hydrogen acceptors
- Do not require aqueous solution
- ACID-BASE CONJUGATE PAIRS (each member within the pair differs from the other one hydrogen)

Buffers



A buffer is mixture of a weak acid and its conjugate base or a weak base and its conjugate acid. Buffers work by reacting with any added acid or base to control the pH

Examples: blood pH, ocean pH

Ocean Acidification: An issue to contemplate

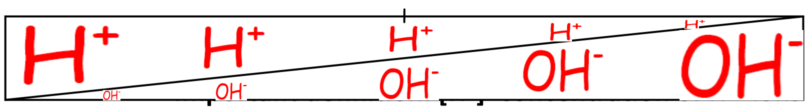
Watch this: NOAA Ocean acidification

Lesson 5: pH and Indicators

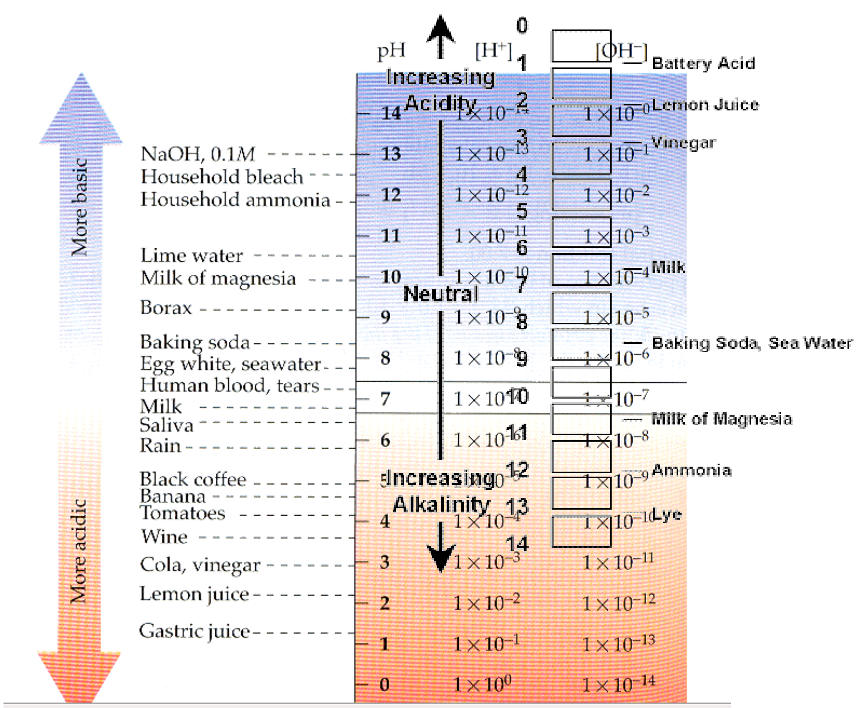
Objective: To categorize substances as acidic or basic based on concentration of H^+ ions in solution (pH)

pH stands for "parts Hydrogen". **(Notice that the H must be capital!)**

Solutions that are **acidic** have pH values **less than 7** in value. **Alkaline** solutions are **basic**, so they have pH values **greater than 7**. An acidic solution has more H^{+1} (hydrogen ions) in it than $(OH)^{-1}$ (hydroxide ions). The opposite is true of basic solutions. **"Neutral"** solutions contain an **equal number** of H^{+1} ions and $(OH)^{-1}$ ions and have a pH value of 7.



- As pH decreases (gets closer to 1), the concentration of H+ increases
- As pH changes by one unit, $[H^+]$ changes by a factor of 10 (pH is a logarithmic scale)
- A pH of 5 compared to a pH of 6: A pH of 5 is 10 times more acidic.
- A pH of 8 compared to a pH of 5: A pH of 5 is 1000 times (10 x 10 x 10) more acidic.



Summary

- pH shows how acidic or alkaline a solution is
- lower the pH the greater the [H+]
- higher the pH the lower the [H+]
- In an acidic substance **$[H^+] > [OH^-]$**
- In a basic substance **$[H^+] < [OH^-]$**
- In a neutral substance **$[H^+] = [OH^-]$**

- The pH scale also measures the strength of acid or base (which is dependent on their ability to IONIZE/produce H^+ ions in solution)
- Logarithmic-based on 10
- SINGLE pH UNIT signifies A TENFOLD CHANGE IN $[H^+]$ CONCENTRATION

The **properties** that an acid displays are due to the presence of H^{+1} ions, and the properties of bases are due to $(OH)^{-1}$ ions. There are different categories of acids and bases **depending on how many H^{+1} or $(OH)^{-1}$ ions are present in solution. We use indicators to measure the level of H^{+1} ions in a solution.**

Acid-base indicators are molecules that are natural dyes, and these dyes happen to be sensitive to pH changes, such that they take on different colors depending on the pH.
If you look at **Table M** in your Reference Tables, you will notice some indicators listed.
Look at bromothymol blue. Read the info like this… if the indicator is put into a solution and it turns yellow, then we know the pH is 6 or less. If it turns blue, the pH is 7.6 or higher. Some indicators have an in-between color as well. Bromothyml blue turns green if the pH is between 6 and 7.6.

Table M
Common Acid–Base Indicators

Indicator	Approximate pH Range for Color Change	Color Change
methyl orange	3.2–4.4	red to yellow
bromthymol blue	6.0–7.6	yellow to blue
phenolphthalein	8.2–10	colorless to pink
litmus	5.5–8.2	red to blue
bromcresol green	3.8–5.4	yellow to blue
thymol blue	8.0–9.6	yellow to blue

Lesson 6: Neutralization and titrations

Objective: To define neutralization as the reaction between an acid and base creating a salt and water.

When an acid **neutralizes** a base, the products are always a **salt** and **water**. If you think about this, it all makes sense… the thing that makes a solution acidic is the presence of **H^{+1}** ions, and the thing that makes a solution a base is the presence of **$(OH)^{-1}$** ions. Therefore, when these two types of solutions are mixed, these two oppositely charged ions attract and bond together to make **water** molecules. Since there are no longer any acid ions or base ions around, the **pH would be 7, or NEUTRAL**, hence being described as a **neutralization** reaction!
A neutralization reaction is a double replacement reaction. You should be able to write a reaction if you know the acid and base involved, since, as mentioned above, a salt and water are always formed.

For example:	$Na(OH)$ **base**	+	$H(NO_3)$ **acid**	→	$Na(NO_3)$ **salt**	+	$H(OH)$ **water**

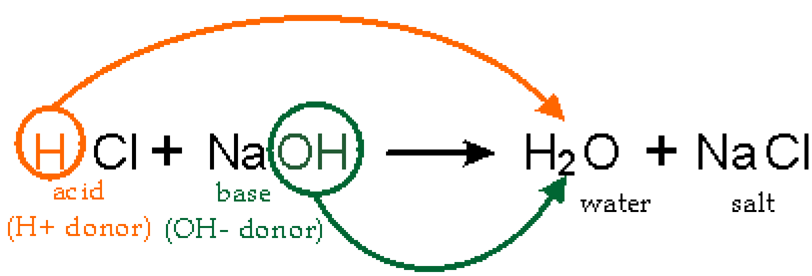
To determine if a reaction is neutralization, look to see if an acid and base is reacting in a double replacement reaction to produce salt and water.

ACID + BASE → SALT + WATER

When an acid reacts with a base, a salt and water are formed. A salt is the general name for an ionic compound.

$HCl + NaOH \rightarrow H_2O + NaCl$
$HBr + KOH \rightarrow H_2O + KBr$
$HNO_3 + NaOH \rightarrow H_2O + NaNO_3$
$H_2SO_4 + 2\ KOH \rightarrow 2\ H_2O + K_2SO_4$ *(note the diprotic acid)*
$2\ HNO_3 + Mg(OH)_2 \rightarrow 2\ H_2O + Mg(NO_3)_2$ *(note the dihydroxy base)*

A neutral solution is formed when the right number of moles of strong acid reacts with strong base.



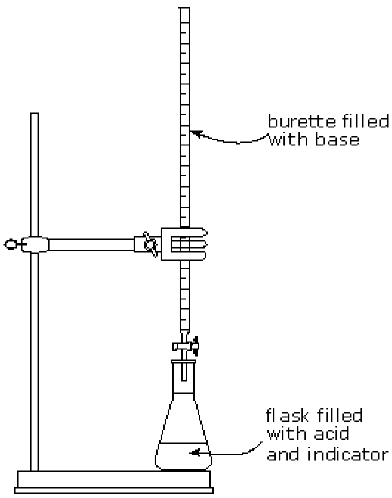
- Neither acidic nor basic; equal concentrations of H+ and OH- (their charges cancel each other out!)
- Double replacement reaction
- H^+ from acid and OH^- from base from water, HOH (H_2O)
- Equal moles of hydrogen ions and hydroxide ions are mixed producing water and a salt
- Exothermic reaction

Titration Reactions

Titrations are used to **CALCULATE THE CONCENTRATION (MOLARITY) OF AN UNKNOWN SOLUTION**; a **laboratory process** in which unknown solution is systematically reacted with a solution of known concentration by adding measured volumes of an acid or a base (of known concentration) to the unknown until **NEUTRALIZATION** occurs. In all neutralization reactions there must be a **1:1 RATIO** between the **MOLES OF H^+ IONS** and the **MOLES OF OH^- IONS**

***The volume can be recorded in mL, L, drops, or any other measurable unit!!**

So, in a titration, when: **$[H^+]$ = $[OH^-]$** or the **MOLES OF H^+ IONS = MOLES OF OH^- IONS**



You've reached the **EQUIVALENCE POINT** of the reaction; this is when the titration is complete (and all data has been retrieved); it's also the **ENDPOINT** if the solution changes color permanently.

Titration formula (Table T): $\mathbf{M_A V_A = M_B V_B}$

Where: $\mathbf{M_A}$ = molarity of acid (H^+)
$\mathbf{V_A}$ = volume of acid
$\mathbf{M_B}$ = molarity of base (OH^-)
$\mathbf{V_B}$ = volume of base

- Use this formula when you are dealing with a titration or neutralization word problem
- Make sure that all units are in agreement when plugging into formula (so they cancel out and you get the right answer!)

Sample Problem 1: What is the concentration of a solution of HI if 0.3 L is neutralized by 0.6 L of 0.2 M solution of KOH?

M_A = ?
V_A = 0.3 L
M_B = 0.2 M
V_B = 0.6 L

$$M_A V_A = M_B V_B$$
$$M_A (0.3L) = (0.2M)(0.6L)$$
$$M_A = \frac{(0.2)(0.6)}{(0.3)}$$
$$M_A = \boxed{0.4M \text{ or moles/L}}$$

Sample Problem 2: A particular acid has an H+ concentration of 0.1 M and a volume of 100 mL. What volume of a base with a 0.5 M [OH-] will be required to neutralize the reaction?

M_A = 0.1M
V_A = 100 mL = 0.100 L
M_B = 0.5 M
V_B = ?

$$M_A V_A = M_B V_B$$
$$(0.1M)(0.100L) = (0.5M)V_B$$
$$V_B = \frac{(0.1)(0.100)}{(0.5)}$$
$$V_B = \boxed{0.02 \text{ L or } 20 \text{ mL}}$$

****Sample Problem 3:** You have 50 mL of 1.0 M $H_2SO_{4(aq)}$. What volume of 0.5 M NaOH would be required to neutralize the acid? Remember → Diprotic Acids yield 2 H^+ ions in solution! You must account for this with the math.

M_A = 2(1.0 M)
V_A = 50 mL
M_B = 0.5 M
V_B = ?

$$2(M_A\ V_A) = M_B V_B$$
$$(2)(1.0M)(50mL) = (0.5M)V_B$$
$$V_B = \frac{(2)(1.0)(50)}{(0.5)}$$
$$V_B = \boxed{200\ mL}$$

Summary

- Titrations are used to calculate the concentration (molarity) of an unknown solution.
- Equivalence point occurs when neutralization is obtained.
- Uses indicators
- There must be a 1:1 ratio between the moles of H^+ IONS and the moles of OH^- ions
- **$[H^+]$ = $[OH^-]$** or the **MOLES OF H^+ IONS = MOLES OF OH^- IONS**

<u>Enrichment:</u> Naming Acids and Bases

<u>Binary Acids</u> → acids that START WITH H and are attached to a NONMETAL; 2 TOTAL ELEMENTS in acid's formula

1st word in name → Begin with *hydro-* & follow it up immediately with the name of the other element while replacing the ending with *-ic*

2nd word in name → add the word acid as the second (last) word in the name

Ex: HCl Hydrochloric acid HI Hydroiodic acid HBr Hydrobromic acid

<u>Ternary Acids</u> → acids that START WITH H & are attached to a POLYATOMIC ION; 3 TOTAL ELEMENTS in the compound formula—USE TABLE E!

1st word in name: There is NO *"HYDRO"* in front! Same rules as the binary acids, except the *–ate/-ite* are replaced by *–ic* and *–ous* (respectively)

So: If the polyatomic ion ends in -ATE it gets replaced by –IC
(I ate in the café and went ic!)
If the polyatomic ion ends in -ITE it gets replaced by –OUS

2nd word in name: Add the word acid as the second word in the name

Ex: HNO_3 Nitric acid Ex: HNO_2 Nitrous acid

<u>Bases</u>

1) Name the 1ST ION in the formula "AS IS." This will always be the + ion and will either be a metal (ex: Na) or a polyatomic ion (ex: NH_4)
2) Name the -OH ON THE END AS HYDROXIDE (separate second word in the name) –
3) Table E says that OH is called "hydroxide"
4) NOTE: it does not matter how many you have of each when it comes to the positive or negative ion (see second example below). Don't ever use prefixes!

Ex: LiOH **Lithium hydroxide** $Mg(OH)_2$ **Magnesium hydroxide**
NH_4OH **Ammonium hydroxide**

<u>CHALLENGE</u>:

A volume of 10 mL of 0.75 M sodium hydroxide neutralizes a 30 mL sample of hypochlorous acid.
Write a balanced equation for the reaction and calculate the concentration of the acid.