

CHEMISTRY NOTES – CHAPTERS 20 AND 21

Acids and Bases - Neutralization

Goals : To gain an understanding of :

1. The properties of acids and bases
2. pH and pOH calculations
3. Definitions of acids and bases
4. Neutralization reactions

NOTES:

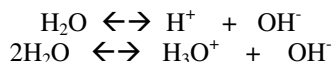
Properties of acids :

- Form hydrogen ions (H^+ or protons) in solution
- Form hydronium (a water molecule bonded to a hydrogen ion shown as H_3O^+ or $H^+ \cdot H_2O$) ions in aqueous solution
- Have a low pH (below 7)
- Neutralize bases to form water and a salt
- Have a sour taste
- Cause indicators to change color
- Are electrolytes (are ionic in nature)
- React with many metals to produce hydrogen gas (single replacement reaction)

Properties of bases :

- Form hydroxyl ions (OH^-) in solution
- Have a high pH (above 7)
- Neutralize acids to form water and a salt
- Have a bitter taste
- Feel slippery
- Cause indicators to change color
- Are electrolytes (are ionic in nature)

Water is neutral, not because it does not have any hydronium (acid) or hydroxyl ions (base) present, but because these ions are equal in concentration (10^{-7} M). They are equal because pure water, even though it is covalently bonded, has some ionic character and will self ionize to form these ions according to either of the following equations.



Note : I will treat the hydrogen ion (H^+) and the hydronium ion (H_3O^+) as the same thing throughout these notes.

Neutral solutions have a pH of 7 and have equal concentrations of hydronium and hydroxyl ions (10^{-7} M).

The pH of a solution is the negative logarithm of the hydrogen ion concentration ($pH = -\log [H^+]$). The logarithm (base 10) of a number is the number to which ten must be raised to get that number. For example the log of .1 is -1 ($.1 = 10^{-1}$). The pH scale is a scale of 0 to 14 and indicates the measure of H^+ ions present - the amount of acidity of a solution.

[H ⁺]	Exponential Notation	log	pH (-log [H ⁺])	Description
1	1 x 10 ⁰	0	0	strong acid
.1	1 x 10 ⁻¹	-1	1	
.01	1 x 10 ⁻²	-2	2	
.001	1 x 10 ⁻³	-3	3	
.0001	1 x 10 ⁻⁴	-4	4	
.00001	1 x 10 ⁻⁵	-5	5	
.000001	1 x 10 ⁻⁶	-6	6	weak acid
.0000001	1 x 10 ⁻⁷	-7	7	neutral ([H ⁺] = [OH ⁻] = 1.0 x 10 ⁻⁷)
.00000001	1 x 10 ⁻⁸	-8	8	weak base
.000000001	1 x 10 ⁻⁹	-9	9	
.0000000001	1 x 10 ⁻¹⁰	-10	10	
.00000000001	1 x 10 ⁻¹¹	-11	11	
.000000000001	1 x 10 ⁻¹²	-12	12	
.0000000000001	1 x 10 ⁻¹³	-13	13	
.00000000000001	1 x 10 ⁻¹⁴	-14	14	strong base

If the pH is less than 7 the solution is considered to be acidic ([H⁺] > 1.0 x 10⁻⁷ and [OH⁻] < 1.0 x 10⁻⁷) and if the pH is above 7 the solution is considered to be basic, or alkaline ([H⁺] < 1.0 x 10⁻⁷ and [OH⁻] > 1.0 x 10⁻⁷).

If the hydrogen ion (H⁺) ion increases the hydroxide ion (OH⁻) will decrease and vice versa.

The ion product constant for water (K_w) at 25 °C is 1.0 x 10⁻¹⁴ (mol/L)². This comes from the equation for the self ionization of water shown above.

$$K_w = [H^+][OH^-] = 1.0 \times 10^{-14}(\text{mol/L})^2 = (1.0 \times 10^{-7} \text{ mol/L})(1.0 \times 10^{-7} \text{ mol/L}) = 1.0 \times 10^{-14}(\text{mol/L})^2$$

This equation can be used to calculate [H⁺] or [OH⁻] as follows:

What is the [H⁺] of an aqueous solution at 25 °C if the [OH⁻] is 2.56 x 10⁻⁵ ?

Rearranging $K_w = [H^+][OH^-]$ to solve for [H⁺] we get $[H^+] = \frac{K_w}{[OH^-]}$

$$[H^+] = \frac{1.0 \times 10^{-14} \text{ (mol/L)}^2}{2.56 \times 10^{-5} \text{ (mol/L)}} = 3.9 \times 10^{-10} \text{ mol/L}$$

The equation to solve for pH from [H⁺] is pH = -log [H⁺]. Here is an example.

What is the pH of a solution in which the [H⁺] is 1.12 x 10⁻⁴ mol/L?

$$\begin{aligned} \text{pH} &= -\log [H^+] \\ &= -\log (1.12 \times 10^{-4}) \\ &= 3.95 \end{aligned}$$

The equation to solve for $[H^+]$ from pH is $[H^+] = 10^{-pH}$. Here is an example.

What is the $[H^+]$ of a solution which has a pH of 6.5 ?

$$\begin{aligned} [H^+] &= 10^{-pH} \\ &= 10^{-6.5} \\ &= 3.2 \times 10^{-7} \text{ mol/L} \end{aligned}$$

Another equation to solve for pH (or $[H^+]$) can be derived from the equation for the ion product constant of water (K_w). Algebraically we can take the log of both sides of the equation for K_w to get the equation :

$$pK_w = pH + pOH$$

$pK_w = 14.00$ therefore the above equation can be rewritten as

$$14.00 = pH + pOH$$

which can be a handy equation to solve for pH from pOH or vice versa. here is an example starting from the hydroxide ion concentration.

What is the pH of a solution which has a $[OH^-]$ ion concentration of 3.0×10^{-6} mol/L ?

$$pOH = -\log[OH^-] = -\log(3.0 \times 10^{-6}) = 5.5$$

Rearranging $14.00 = pH + pOH$ to solve for pH we get $pH = 14.00 - pOH$

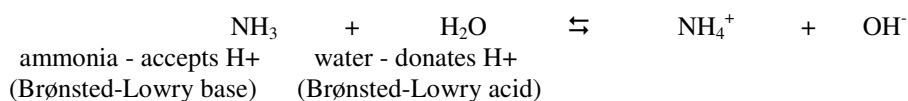
$$\text{solving : } pH = 14.00 - 5.5 = 8.5$$

Svante Arrhenius (Sweden, 1887) proposed that acids are substance which yield hydrogen ions (H^+) in aqueous solution and bases are substances which yield hydroxide (OH^-) ions in aqueous solution. Examples of Arrhenius acids are hydrochloric acid (HCl), sulfuric acid (H_2SO_4), and phosphoric acid (H_3PO_4). Examples of Arrhenius bases are sodium hydroxide (NaOH), calcium hydroxide ($Ca(OH)_2$), and aluminum hydroxide ($Al(OH)_3$).

Monoprotic acids have only one ionizable hydrogen ion (proton). Examples are hydrochloric acid (HCl) and nitric acid (HNO_3). Diprotic acids have two ionizable hydrogen ions. Examples are sulfuric acid (H_2SO_4) and carbonic acid (H_2CO_3). Triprotic acids have three ionizable hydrogen ions. An example is phosphoric acid (H_3PO_4).

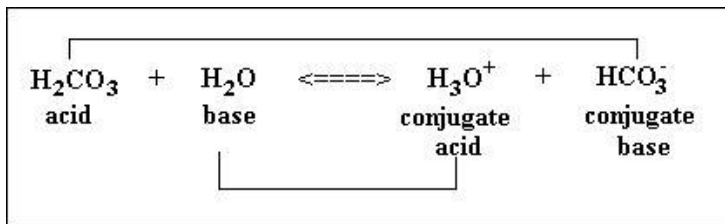
Only the hydrogens involved in very polar or ionic bonds will ionize to form ions. These are hydrogens that are bonded to highly electronegative elements found in the upper right hand side of the periodic table. Ionizable hydrogens are more electropositive and are generally written as the first element in the formula of a compound. For example in the formula for acetic acid ($HC_2H_3O_2$) only the first hydrogen (in blue) is ionizable.

The Brønsted-Lowry definitions of acids and bases are that acids are proton (hydrogen ion) donors and bases are proton acceptors. All Arrhenius acids and bases are included in the definitions of Brønsted-Lowry definitions, but there are some Bronsted-Lowry acids or bases that are not Arrhenius acids or bases. For example ammonia (NH_3) is not an Arrhenius base - it does not form OH^- ions (in the equation below water provides the OH^- ion). It is a Brønsted-Lowry base because it will accept a hydrogen ion as shown in the following equation.



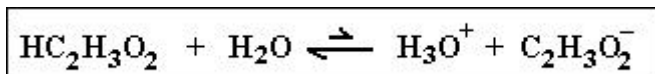
A conjugate acid is the substance formed when a Brønsted-Lowry base accepts a hydrogen ion. A conjugate base is the substance formed when a Brønsted-Lowry acid donates a hydrogen ion. In the above equation the ammonium ion (NH_4^+) is the conjugate acid formed by the Brønsted-Lowry base ammonia (NH_3) accepting a hydrogen ion. The hydroxide ion (OH^-) is the conjugate base formed by the Brønsted-Lowry acid water donating a hydrogen ion. Conjugate acids and bases are always written on the

product side of a chemical equation. Conjugate acids behave just as a Brønsted-Lowry acid in the reverse direction (donate hydrogen ions) and conjugate bases behave just as Brønsted-Lowry bases in the reverse direction. Here is another example.



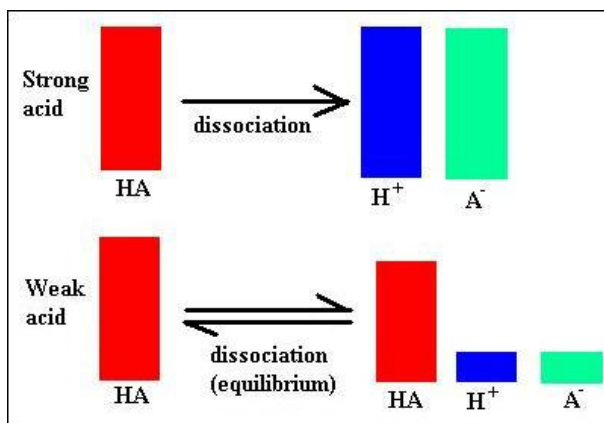
Amphoteric substances are substances which can act as both an acid and a base. An example is water. Water when it self ionizes forms a hydrogen ion (acid) and a hydroxide ion (base).

Acids and bases can be classified as strong or weak depending on the extent to which they ionize. Strong acids and bases ionize to a large extent to form many hydrogen or hydroxide ions and have a large acid or base dissociation constant (K_a or K_b). Weak acids or bases ionize to a small extent to form few hydrogen or hydroxide ions and have a small K_a or K_b . An example of a weak acid is acetic acid. It will form few hydrogen ions and most of the acid stays in the undissociated form as shown in the equation below.



The larger arrow in the reverse direction indicates that the equilibrium position lies to the left indicating more undissociated acid and fewer hydronium ions being formed.

Here is another way to picture the difference between a strong and weak acid. HA is the undissociated acid, H^+ indicates the dissociated hydrogen ions and A^- indicates the dissociated anion of the acid.

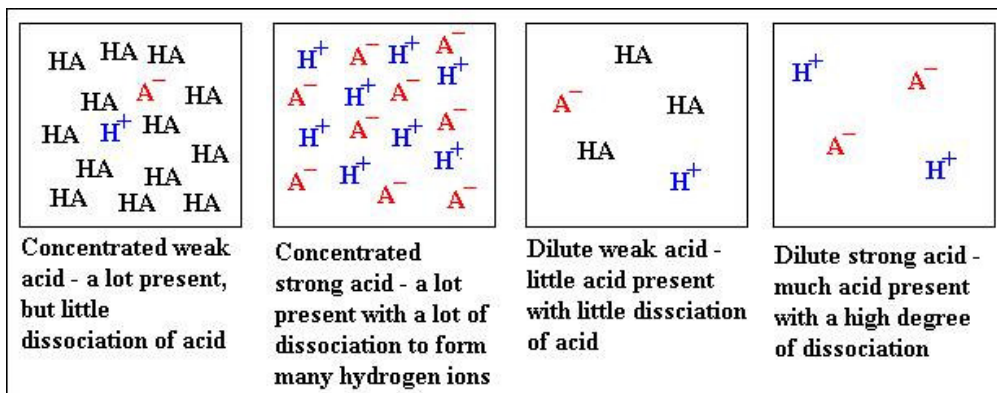


Diprotic and triprotic acids dissociate stepwise, meaning all the hydrogen ions do not dissociate simultaneously. For example phosphoric acid (H_3PO_4) would dissociate in three steps :

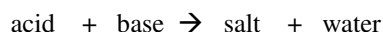
- $\text{H}_3\text{PO}_4 \rightleftharpoons \text{H}^+ + \text{H}_2\text{PO}_4^-$
- $\text{H}_2\text{PO}_4^- \rightleftharpoons \text{H}^+ + \text{HPO}_4^{2-}$
- $\text{HPO}_4^{2-} \rightleftharpoons \text{H}^+ + \text{PO}_4^{3-}$

Step a would be the have the largest K_a . Each successive step would have a smaller K_a as it becomes increasingly difficult to remove a positive hydrogen ion from a more negative anion.

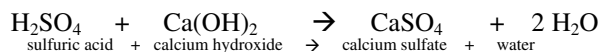
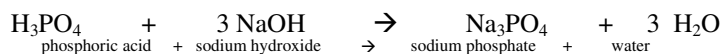
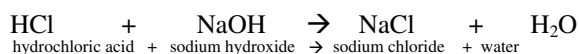
The terms concentrated and dilute should not be confused with the terms strong and weak in relation to acids and bases. Concentrated and dilute refer to how much acid or bases is present, not the extent to which that acid or base will ionize. The following diagrams should illustrate the differences. HA refers to the undissociated acid, H^+ indicates the hydrogen ions and A^- indicates the anion of the acid.



Neutralization reactions are reactions between acid and bases. In aqueous solution neutralization reactions produce a salt and water. The generalized equation is thus :



Here are some specific examples :



The term "neutralization" can be misleading. Not all neutralization reactions result in a neutral solution. Only neutralization reactions which have equivalent amounts of hydrogen ions and hydroxide ions from acids and bases will result in neutral solutions. The balanced equation and a little bit of stoichiometry can be used to calculate the equivalent number of moles needed. Here is an example of such a calculation :

(1) Write and balance the equation for the neutralization of carbonic acid by lithium hydroxide and then (2) determine the number of moles of lithium hydroxide needed to neutralize 1.12 moles of carbonic acid.



2. Using the mole ratio from the balanced equation :

$$1.12 \text{ mol } \cancel{\text{H}_2\text{CO}_3} \left(\frac{2 \text{ mol LiOH}}{1 \text{ mol } \cancel{\text{H}_2\text{CO}_3}} \right) = 2.24 \text{ mol LiOH}$$

Titration is a powerful analytical tool used to determine the concentration of an unknown from the concentration of a known (standard) solution. Four steps are needed to perform the calculations of a titration between an acid and a base.

1. Write and balance the neutralization reaction.
2. Determine the number of moles used of the standard (known) solution. (moles = M x V)
3. Determine the number of moles present in the unknown using the mole ratio from the balanced equation.
4. Determine the molarity (moles/L) of the unknown from step 3 and the volume of the unknown given in the problem.

Here is an example.

What is the concentration of a 50.0 mL (=0.0500L) solution of HCl which is neutralized by 25.0 mL (=0.0250L) of a 0.50 M (mol/L) solution of Ca(OH)₂?

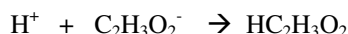


$$2. \cancel{.025 \text{ L}} \left(\frac{0.50 \text{ mol Ca(OH)}_2}{\cancel{\text{L}}} \right) = .013 \text{ mol Ca(OH)}_2$$

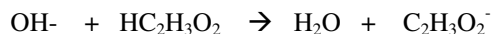
$$3. \cancel{.013 \text{ mol Ca(OH)}_2} \left(\frac{2 \text{ mol HCl}}{1 \text{ mol Ca(OH)}_2} \right) = .026 \text{ mol HCl}$$

$$4. \frac{.026 \text{ mol HCl}}{.0500 \text{ L}} = .52 \text{ M HCl}$$

Buffers are substances or combinations of substances which keep pH levels fairly constant when small amounts of acid or base are added. Buffers are formed from the combination of a weak acid and one of its salts or a weak base and one of its salts. An example of a buffer is acetic acid (weak) and sodium acetate, a salt of acetic acid. If a small amount of acid is added to a system containing these substances the hydrogen ion from the acid will react with the acetate ion from the sodium acetate to form undissociated acetic acid effectively removing the hydrogen ion to keep pH constant :

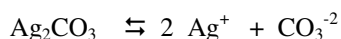


If a base is added the OH⁻ ion from the base will react with the acetic acid to form water and the acetate ion to effectively remove the added OH⁻ ion :



So, whether acid or base are added they are removed to keep pH constant.

The common ion effect is the lowering of the solubility of a substance by the addition of a common ion. This can be easily seen in relation to LeChâtelier's principle. At equilibrium a slightly soluble salt such as silver carbonate will be saturated.



If silver iodide (AgI) is added to the system it will increase the [Ag⁺] - the common ion - and cause the reaction to shift to the left. This is the undissociated, or non dissolved, state indicating the lowering of the solubility of the silver carbonate.

Acids and Bases

Properties of Acids and Bases

I. Acids

A. Properties of Acids

1. Aqueous solutions of acids have a sour taste
2. Acids change the color of acid-base indicators
3. Some acids react with active metals to release hydrogen

$$\text{Zn(s)} + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{ZnSO}_4(\text{aq}) + \text{H}_2(\text{g})$$
4. Acids react with bases to produce salts and water (neutralization)

$$\text{HCl(aq)} + \text{NaOH(aq)} \rightarrow \text{NaCl(aq)} + \text{H}_2\text{O(l)}$$
5. Aqueous solutions of acids conduct electric current (they are electrolytes)

II. Some Common Industrial Acids

A. H₂SO₄ - Sulfuric Acid

1. Used in making fertilizer, paper, petroleum products, car batteries
2. Highest production chemical in the U.S.

B. HNO₃ - Nitric Acid

1. Unstable, volatile liquid in pure state
2. Yellowing caused by the formation of nitrogen dioxide gas

3. Used in making fertilizers, explosives, rubber, plastics, pharmaceuticals
 4. Stains skin and other proteins yellow
- C. H_3PO_4 - Phosphoric Acid
1. Not a common laboratory acid
 2. Flavoring agent in sodas
 3. Used in making fertilizers and detergents
- D. HCl - Hydrochloric Acid
1. HCl in the stomach aids the digestion of proteins
 2. Used in pickling steel, recovering magnesium from sea water, cleaning masonry and correcting pool pH
- E. Acetic Acid
1. Acid component of vinegar
 2. Concentrated "glacial" acetic acid used in making chemicals for plastic manufacturing

III. Bases

A. Properties of Bases

1. Aqueous solutions of bases have a bitter taste
2. Bases change the color of acid-base indicators
3. Dilute aqueous solutions of bases feel slippery
4. Bases react with acids to produce salts and water
5. Aqueous solutions of bases conduct electric current (they are electrolytes)

IV. Arrhenius Acids and Bases Svante Arrhenius, Swedish chemist (1859-1927)

A. Arrhenius Acid

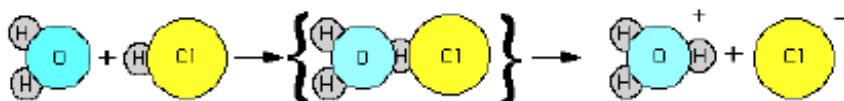
1. A chemical compound that increases the concentration of hydrogen ions, H^+ , in aqueous solution

B. Arrhenius Base

1. A substance that increases the concentration of hydroxide ions, OH^- , in aqueous solution

C. Aqueous solutions of acids

1. Acids are molecular compounds that ionize in solution



D. Strength of Acids

1. Strong acids ionize completely in solution and are strong electrolytes
2. Weak acids ionize only slightly and are weak electrolytes

<i>Strong acids</i>	<i>Weak acids</i>
H_2SO_4	HSO_4^-
HClO_4	H_3PO_4
HCl	HF
HNO_3	CH_3COOH
HBr	H_2CO_3
HI	H_2S
	HCN
	HCO_3^-

Bronsted-Lowry Acids and Bases

I. Bronsted-Lowry Acids and Bases

A. Bronsted-Lowry Acid

1. A molecule or ion that is a proton donor (a proton is a hydrogen ion H^+)

B. Bronsted-Lowry Base

1. A molecule or ion that is a proton acceptor

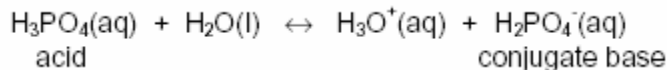
C. Bronsted-Lowry Acid-Base Reaction

1. A reaction in which protons are transferred from the acid to the base

II. Conjugate Acids and Bases

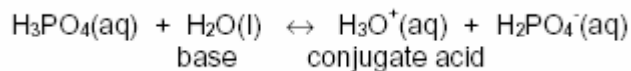
A. Conjugate Base

1. The species that remains after an acid has given up a proton



B. Conjugate Acid

1. The species that is formed when a base gains a proton



III. Amphoteric Compounds

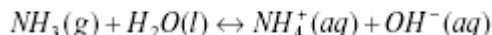
A. Amphoteric

1. Any species that can react as either an acid or a base

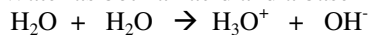
a. Water as a base



b. Water as an acid



c. Water as both an acid and a base



H_3O^+ is called the hydronium ion and is what actually exists in an acidic solution. There are no free H^+ ions

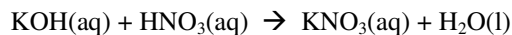
Neutralization Reactions

I. Strong Acid-Strong Base Neutralization

A. Neutralization

1. The reaction of hydronium ions and hydroxide ions to form water molecules

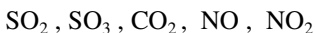
B. Neutralization Rxns



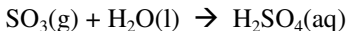
II. Acid Rain

A. Formation of Acid Rain

1. Nonmetallic oxides enter the atmosphere as a result of coal burning, auto exhaust, other forms of air pollution



2. Nonmetallic oxides combine with water to form oxyacids



B. Reactions of Acid Rain

1. Reaction with marble (metamorphic calcium carbonate)

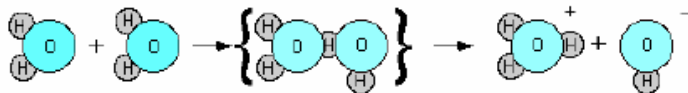
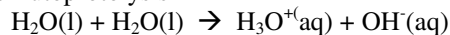


Aqueous Solutions and the Concept of pH

I. Hydronium Ions and Hydroxide Ions

A. Self-Ionization of Water

1. Autoprotolysis



B. Ionization Constant for Water (K_w)

- $K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = (1.0 \times 10^{-7})(1.0 \times 10^{-7}) = 1.0 \times 10^{-14}$
- K_w is a constant at ordinary ranges of room temperatures

C. Neutral, Acidic, and Basic Solutions

1. Neutral

- $[\text{H}_3\text{O}^+] = [\text{OH}^-]$ (pH = 7.00)

2. Acidic

- $[\text{H}_3\text{O}^+] > [\text{OH}^-]$ (pH < 7.00)

3. Basic

- $[\text{H}_3\text{O}^+] < [\text{OH}^-]$ (pH > 7.00)

II. The pH Scale

A. pH

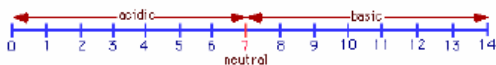
- The negative of the common logarithm of the hydronium ion concentration
 - $\text{pH} = -\log [\text{H}_3\text{O}^+]$

B. pOH

- The negative of the common logarithm of the hydroxide ion concentration
 - $\text{pOH} = -\log [\text{OH}^-]$

$$\text{pH} + \text{pOH} = 14.0$$

C. The pH Scale



III. Calculations Involving pH

A. Calculation of pH from $[\text{H}_3\text{O}^+]$ Concentration

See worksheet

Determining pH and Titrations

I. Indicators and pH strips

A. Acid-Base Indicators

- Compounds whose colors are sensitive to pH

B. Transition Interval

- pH range over which an indicator color change occurs
- Indicators are useful when they change color in a pH range which includes the endpoint of the reaction

C. Choosing Indicators (indicator must change color at the proper pH)

II. Acid-Base Titration

A. Titration

- Controlled addition of the measured amount of a solution of a known concentration required to react completely with a measured amount of solution of unknown concentration

B. Equivalence Point

- The point at which the solutions used in a titration are present in chemically equivalent amounts (equal mols)

C. Titration Curves

1. End point

- the point in a titration at which the reaction is just completed (mol acid = mol base)