Acids and Bases

Chemistry Chapter 18

Objectives

- **Classify** solutions as acidic, basic, or neutral
- **Identify** the physical and chemical properties of acids and bases
- **Compare** the Arrhenius, Brønsted-Lowry, and Lewis models of acids and bases

Acids & Bases

- An **acid** is a compound that when dissolved in water increases the concentration of hydrogen ion, H+
- A **base** is a substance that when dissolved in water increases the concentration of the hydroxide ion, OH--

  NaOH → Na+ + OH-
- Acids and bases are conductors of electricity
- Acids and bases can be indentified by their reactions with some metals and carbonates

Acid	**Base**
General Form: Hx	General Form: xOH
Turns blue litmus red	Turns red litmus blue
Change phenolphthalein from red to colorless	Change phenolphthalein from colorless to pink
taste sour	Taste bitter and feel slippery

Acids & Bases

- Magnesium and zinc react with acids to produce hydrogen gas
- All water solutions contain hydrogen ions (H^+) and hydroxide ions (OH^-)
- An **acidic solution** contains more hydrogen ions than hydroxide ions
- A **basic solution** contains more hydroxide ions than hydrogen ions

Naming Bases

- Bases are named in the same way as other ionic compounds. First name the cation followed by the name of the anion
  - NaOH is sodium hydroxide
  - $Ca(OH)_2$ is calcium hydroxide

Naming Acids

Binary Acids

- All acids start with hydrogen
  - Acid name begins with the prefix **hydro**-
- Binary Acids end with the anion (negative ion)
  - The stem of the anion has the suffix –**ic** and followed by the word acid (remove the –ide ending)

HCl

Hydro-

(stem)-***ic*** acid

Hydrochlor***ic*** acid

Naming Acids

- An **Oxyacid** is an acid that contains both a hydrogen & an oxyanion (polyatomic)
  - Anion ends in –**ite**
    - Acid name is the stem of the anion with the suffix –**ous**, followed by the word acid
  - Anion ends in –**ate**
    - Acid name is the stem of the anion with the suffix –**ic**, followed by the word acid

$\mathbf{H_2SO_3}$

SO_3^{-2} Sulf***ite***

(stem) –***ous*** acid

Sulfur*ous* acid

$\mathbf{H_2SO_4}$

SO_4^{-2} Sulf***ate***

(stem) –***ic*** acid

Sulfur*ic* acid

Acid/Base theories

Three theories have been developed to explain the observed behavior of acids and bases

1. **The Arrhenius Acids and Bases** (1887)
   - **Acids -** hydrogen containing compounds
     - $HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$
     - yields H^+ ions
   - **Bases** - hydroxide containing compounds
     - $NaOH_{(aq)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$
     - yields OH^- ions

Bronsted-Lowry Acids and Bases

2. <u>Bronsted-Lowry Acids and Bases (1923)</u>

- This model is a more inclusive model of acids and bases
- B-L theory states:
  - An acid as a hydrogen-ion donor
  - A base as a hydrogen-ion acceptor

$$HCl_{(aq)} + H_2O_{(l)} \longleftrightarrow H_3O^+_{(l)} + Cl^-_{(aq)}$$

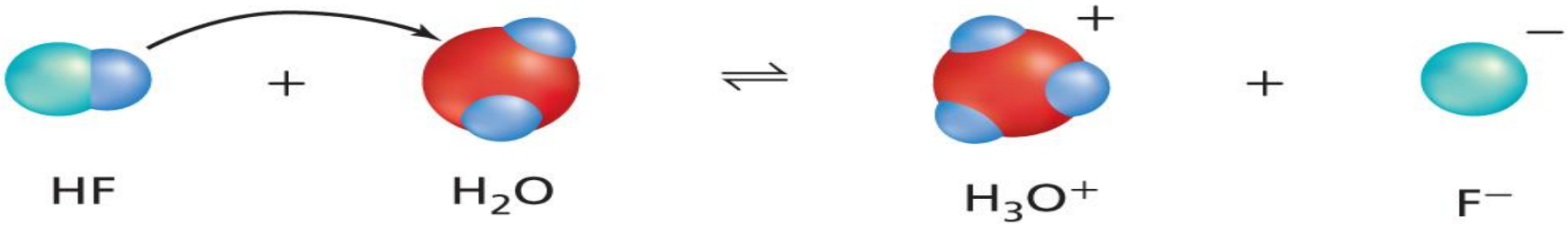
Acid **Base** Conjugate Acid Conjugate Base

- A Conjugate acid is the particle produced when a base gains a hydrogen ion

- A Conjugate base is the particle produced when an acid donates a hydrogen ion

Bronsted-Lowry Acids and Bases

- A **conjugate** acid-base pair consists of two substances related to each other by donating and accepting a single hydrogen ion.



Acid **Base** Conjugate Acid Conjugate Base

$$NH_{3\,(aq)} + H_2O_{\,(l)} \longleftrightarrow OH^-_{\,(aq)} + NH_4^+{}_{(aq)}$$

Base **Acid** Conjugate Base Conjugate Acid

- **Amphoteric** – a substance that can act as both an acid & base
  - Example - water

The Lewis Model

3. **The Lewis model** includes all the substances classified as Brosted-Lowry acids and bases and MANY more

- **Lewis acid** – a substance that can **accept** a pair of electrons to form a covalent bond
- **Lewis base** – a substance that can **donate** a pair of electrons to form a covalent bond

Table 18.2 Three Models for Acids and Bases

Model	Acid Definition	Base Definition
Arrhenius	H^+ producer	OH^- producer
Brønsted-Lowry	H^+ donor	H^+ acceptor
Lewis	electron-pair acceptor	electron-pair donor

Monoprotic and Polyprotic Acids

- An acid that can donate only one hydrogen ion is **monoprotic acid**
  - Example: HNO_3

- Acids that can donate more than one hydrogen are **polyprotic acids**
  - Diprotic acids – 2 ionizable hydrogens
    - H_2SO_4
  - Triprotic acids – 3 ionizable hydrogens
    - $H_3C_6H_5O_7$ (citric acid)

Objectives Section 18.2

- **Compare** the strength of a weak acid with the strength of its conjugate base.
- **Explain** the relationship between the strengths of acids and bases and the values of their ionization constants.
- **Relate** the strength of an acid or base to its degree of ionization.

Strengths of Acids

- The strength of an acid or base is determined by the degree of ionization
- Acids that ionize completely are strong acids
  - Because they produce the maximum number of hydrogen ions, strong acids are good conductors of electricity
  - With a strong acid, the conjugate base is a weak base
- Acids that ionize only partially in dilute aqueous solutions are called weak acids

Table 18.3 Ionization Equations

Strong Acids		Weak Acids	
Name	Ionization Equation	Name	Ionization Equations
Hydrochloric	$HCl \rightarrow H^+ + Cl^-$	Hydrofluoric	$HF \rightleftharpoons H^+ + F^-$
Hydroiodic	$HI \rightarrow H^+ + I^-$	Acetic	$HC_2H_3O_2 \rightleftharpoons H^+ + C_2H_3O_2^-$
Perchloric	$HClO_4 \rightarrow H^+ + ClO_4^-$	Hydrosulfuric	$H_2S \rightleftharpoons H^+ + HS^-$
Nitric	$HNO_3 \rightarrow H^+ + NO_3^-$	Carbonic	$H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$
Sulfuric	$H_2SO_4 \rightarrow H^+ + HSO_4^-$	Hypochlorous	$HClO \rightleftharpoons H^+ + ClO^-$

Strength of Acids

- **The equilibrium constant**, K_{eq}, provides a quantitative measure of the degree of ionization of an acid
- The **acid ionization constant** is the value of the equilibrium constant expression for the ionization of a weak acid, K_a.
  - Quantitative measure of acids strength
  - K_a indicates whether products or reactants are favored at equilibrium

Table 18.4 Ionization Constants for Weak Acids

Acid	Ionization Equation	K_a (298 K)
Hydrosulfuric, first ionization	$H_2S \rightleftharpoons H^+ + HS^-$	8.9×10^{-8}
Hydrosulfuric, second ionization	$HS^- \rightleftharpoons H^+ + S^{2-}$	1×10^{-19}
Hydrofluoric	$HF \rightleftharpoons H^+ + F^-$	6.3×10^{-4}
Hydrocyanic	$HCN \rightleftharpoons H^+ + CN^-$	6.2×10^{-10}
Acetic	$CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$	1.8×10^{-5}
Carbonic, first ionization	$H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$	4.5×10^{-7}
Carbonic, second ionization	$HCO_3^- \rightleftharpoons H^+ + CO_3^{2-}$	4.7×10^{-11}

Strength of Bases

- A base that dissociates completely into metal ions and hydroxide ions is known as a **strong base**
  - A weak base ionizes only partially in dilute aqueous solution
- The **base ionization constant,** K_b, is the value of the equilibrium constant expression for the ionization of a base

Table 18.5 Dissociation Equations for Strong Bases

$NaOH(s) \rightarrow Na^+(aq) + OH^-(aq)$
$KOH(s) \rightarrow K^+(aq) + OH^-(aq)$
$RbOH(s) \rightarrow Rb^+(aq) + OH^-(aq)$
$CsOH(s) \rightarrow Cs^+(aq) + OH^-(aq)$
$Ca(OH)_2(s) \rightarrow Ca^{2+}(aq) + 2OH^-(aq)$
$Ba(OH)_2(s) \rightarrow Ba^{2+}(aq) + 2OH^-(aq)$

Table 18.6 Ionization Constants of Weak Bases

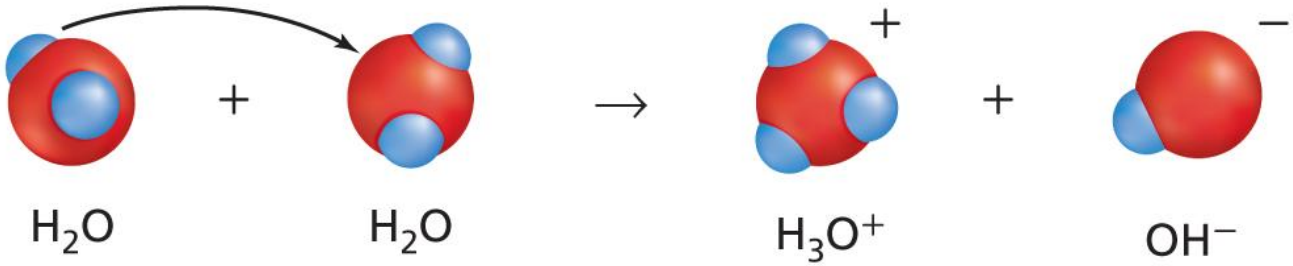
Base	Ionization Equation	K_b (298 K)
Ethylamine	$C_2H_5NH_2(aq) + H_2O(l) \rightleftharpoons C_2H_5NH_3^+(aq) + OH^-(aq)$	5.0×10^{-4}
Methylamine	$CH_3NH_2(aq) + H_2O(l) \rightleftharpoons CH_3NH_3^+(aq) + OH^-(aq)$	4.3×10^{-4}
Ammonia	$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$	2.5×10^{-5}
Aniline	$C_6H_5NH_2(aq) + H_2O(l) \rightleftharpoons C_6H_5NH_3^+(aq) + OH^-(aq)$	4.3×10^{-10}

Objectives Section 18.3

- **Explain** pH and pOH
- **Calculate** the pH and pOH of aqueous solutions
- **Main idea**: pH and pOH are logarithmic scales that express the concentrations of hydrogen ions and hydroxide ions in aqueous solutions

Ion Product constant for water

- Pure water contains equal concentrations of H^+ and OH^- ions



- The **ion product constant for water** is the value of the equilibrium constant expression for the self-ionization of water
- The ion production of water, $K_w = [H^+][OH^-]$
  - [ ] means concentration
- With pure water both $[H^+]$ & $[OH^-]$ are equal to $1.0 \times 10^{-7} M$

Ion Product constant for water

- $[H^+][OH^-] = 10^{-7} * 10^{-7} = 10^{-14}$ M
  - This is a neutral solution
  - $K_{water} = 1.0 \times 10^{-14}$
  - If $[H^+]$ goes up, $[OH^-]$ must go down and vice versa

- Concentrations of H^+ ions are often small numbers expressed in exponential notation
  - Example: $3.5 \text{ X } 10^{-9}$ M

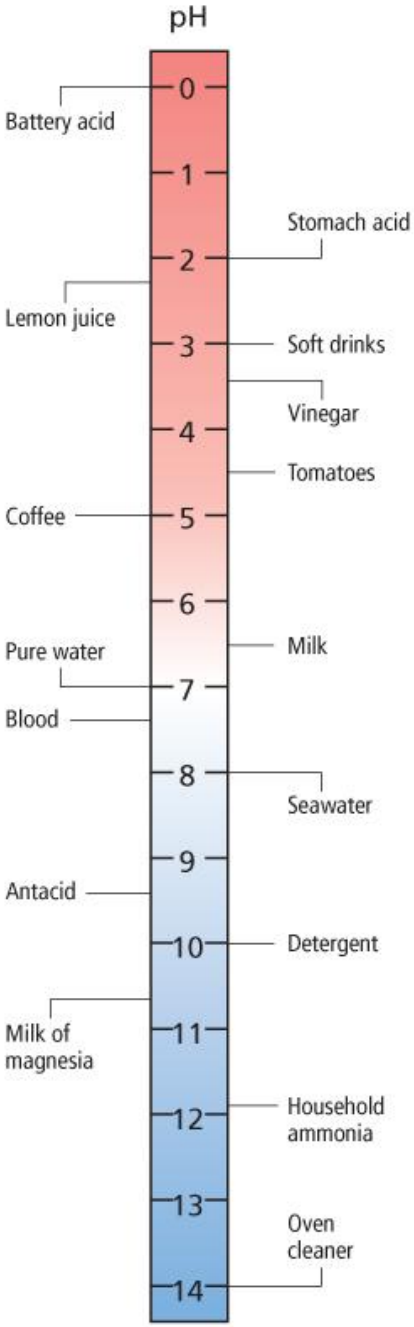
- **pH** is the negative logarithm of the hydrogen ion concentration of a solution
  - $pH = -\log [H^+]$

pH and pOH

The pH scale

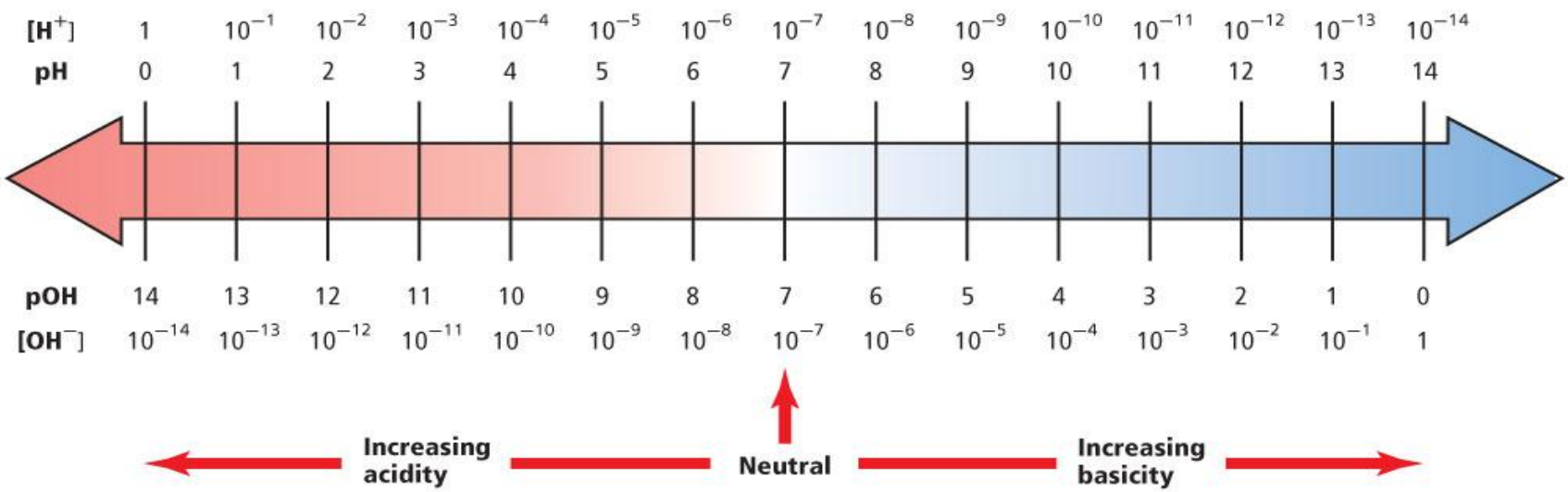
- pH stands for 'potential of hydrogen' or hydrogen ion concentrations
  - $pH = -\log [H^+]$

- pH scale is from 0 to 14
  - **Acidic solution**: pH < 7.0
    - $[H^+]$ greater than $1x10^{-7}$M
  - Neutral solution: pH = 7.0
  - **Basic solution:** pH > 7.0
    - $[H^+]$ less than $1x10^{-7}$ M



pH and pOH

- **pOH** of a solution is the negative logarithm of the hydroxide ion concentration
  - $pOH = -\log [OH^-]$
  - The sum of pH and pOH equals 14



Calculating pH and POH

pH Formulas

$pH = -\log [H^+]$

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

$pH + pOH = 14$

What is the pH of a solution with a $[H^+]$ of 1.0×10^{-10}M?

$pH = -\log [H^+]$

$= -\log (1.0 \times 10^{-10})$

$= -(-10)$

$pH = 10.$

The pH of an unknown solution is 6.00. What is the hydrogen-ion concentration?

$pH = -\log [H^+]$

$6.00 = -\log [H^+]$

$-6.00 = \log [H^+]$ (use the 10x button on your calculator to get antilog)

$[H^+] = 1.00 \times 10^{-6}$ M

Calculating pH and POH

pH Formulas

$pH = -\log [H^+]$

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

$pH + pOH = 14$

What is the pH of a solution if $[OH^-] = 4.00 \times 10^{-11}$ M?

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

$pOH = -\log (4.00 \times 10^{-11})$

$= -(-10.4)$

$pOH = 10.4$

$pH + pOH = 14$

$pH = 14 - pOH$

$pH = 14 - 10.4$

$pH = 3.6$

What is the pH of a 2.700×10^{-2} M KOH solution?

$pOH = -\log [OH-]$

$= -\log (2.700 \times 10^{-2})$

$= -(-1.569)$

$pOH = 1.569$

$pH + pOH = 14$

$pH = 14 - pOH$

$= 14 - 1.569$

$pH = 12.431$

Objectives Section 18.4

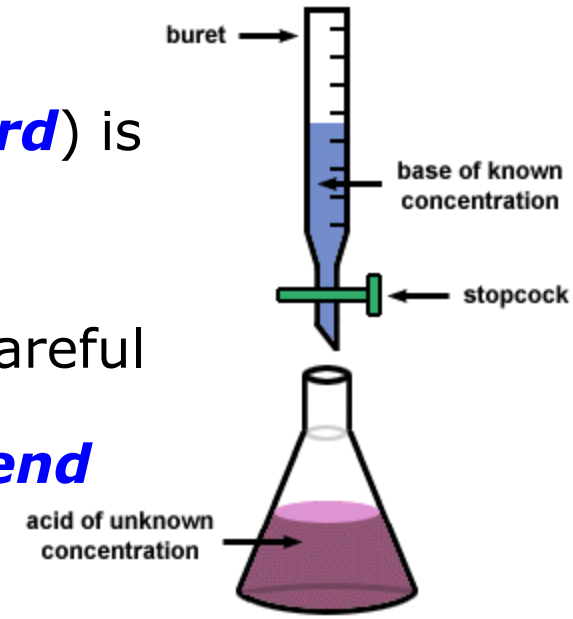
- **Write** chemical equations for neutralization reactions
- **Explain** how neutralization reactions are used in acid-base titrations.
- **Compare** the properties of buffered and unbuffered solutions.

Reactions between Acids and Bases

- A **neutralization reaction** is a reaction in which an acid and a base in an aqueous solution react to produce a salt and water.
  - Acid + Base yield salt + water
  - Neutralization is a double replacement reaction

- **Titration -** method used to determine the concentration of solution (acid or base)

- A soln of known concentration (the ***standard***) is added to a measured amount of the soln of unknown concentration until the soln is completely neutralized
  - ***titrant*** - the known solution added by careful steps using a buret
  - Continues until an indicator signals the ***end point***

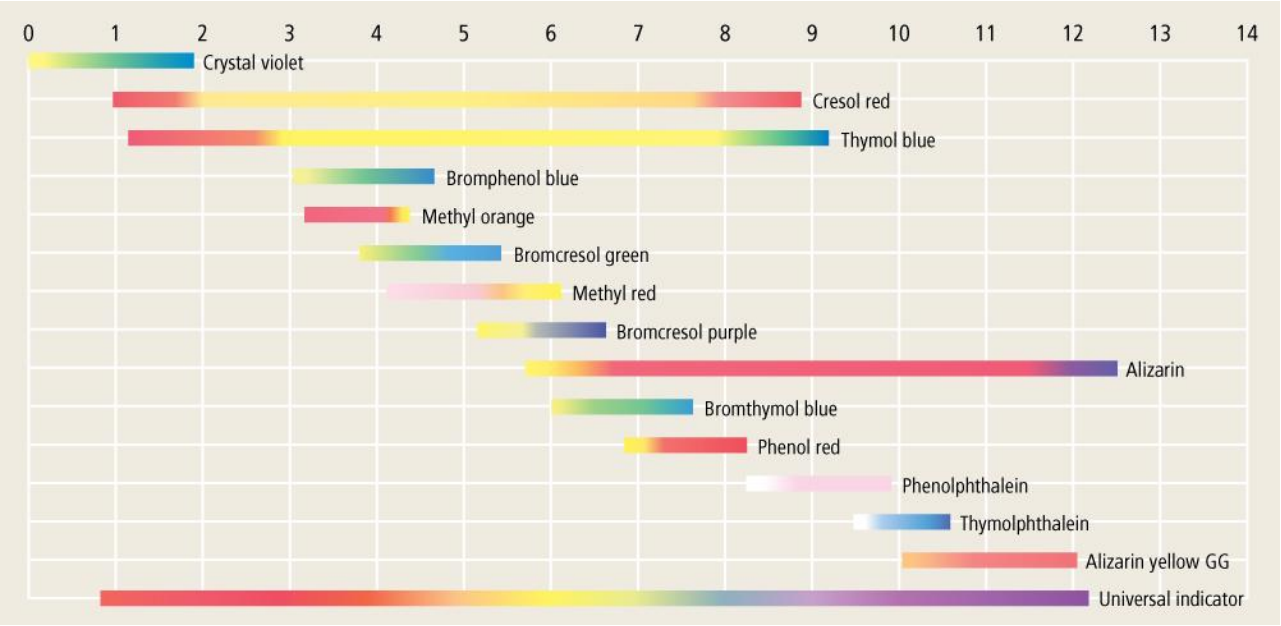


Titration

- **End point** - point in a titration at which neutralization is achieved
  - Indicator changes color
  - End point more technically is called the ***equivalence point***
- **Equivalence point** - the point in a titration at which a stoichiometrically equivalent amount of base has been added to the acid.
  - # of equivalent weights of titrant = # of equivalent weights of unknown
  - Equivalent (equiv) - one equivalent is the amount of an acid (or base) that can give one mole hydrogen (or hydroxide) ions.
  - Normality (N) = equiv
  - $N_1 \times V_1 = N_2 \times V_2$ or $M_A \times V_A = M_B \times V_B$

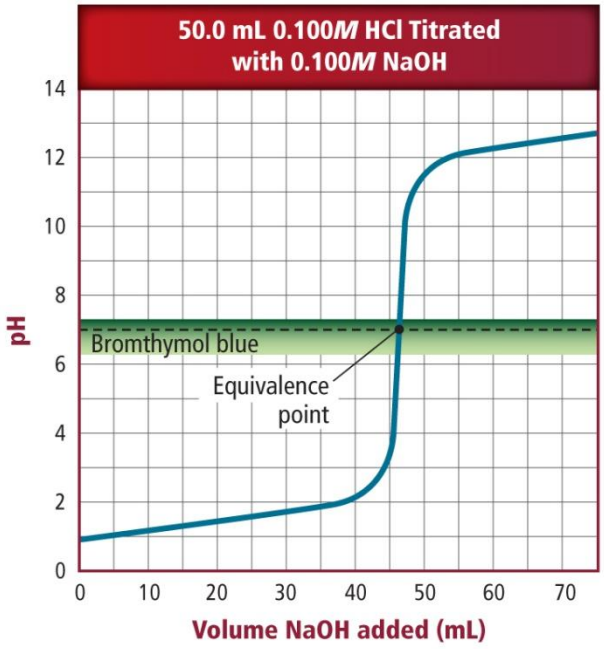
Titration

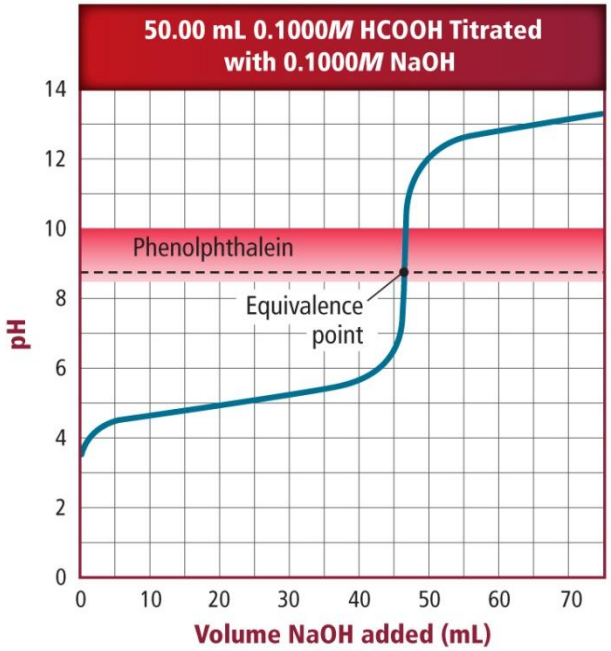
- Chemical dyes whose color are affected by acidic and basic solutions are called **acid-base indicators**
- An **end-point** is the point at which an indicator used in a titration changes color.
- An indicator will change color at the equivalence point.



Steps in a Neutralization reaction

1. A measured volume of acid soln of unknown concentration is added to a flask
2. Several drops of the indicator are added to the soln
3. Measured volumes of a base of **Standard Solution** (known concentration) are mixed into the acid until the indicator just barely changes color





Key concepts

- The concentrations of hydrogen ions and hydroxide ions determine whether an aqueous solution is acidic, basic, or neutral.
  - An Arrhenius acid must contain an ionizable hydrogen atom. An Arrhennius base must contain an ionizable hydroxide group.
  - A Brønsted-Lowry acid is a hydrogen ion donor. A Brønsted-Lowry base is a hydrogen ion acceptor.
  - A Lewis acid accepts an electron pair. A Lewis base donates an electron pair.
- Strong acids and strong bases are completely ionized in a dilute aqueous solution. Weak acids and weak bases are partially ionized in a dilute aqueous solution.

Key Concepts

- For weak acids and weak bases, the value of the acid or base ionization constant is a measure of the strength of the acid or base.
- The ion product constant for water, K_w, equals the product of the H^+ ion concentration and the OH^- ion concentration. $K_w = [H^+][OH^-]$
- The pH of a solution is the negative log of the hydrogen ion concentration. The pOH is the negative log of the hydroxide ion concentration. pH plus pOH equals 14.
  $pH = -\log [H^+]$
  $pOH = -\log [OH^-]$
  $pH + pOH = 14.00$
- A neutral solution has a pH of 7.0 and a pOH of 7.0 because the concentrations of hydrogen ions and hydroxide ions are equal

Key concepts

- In a neutralization reaction, an acid and a base react to form a salt and water.
- The net ionic equation for the neutralization of a strong acid by a strong base is $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$
- Titration is the process in which an acid-base neutralization reaction is used to determine the concentration of a solution.
- Buffered solutions contain mixtures of molecules and ions that resist changes in pH.