Aqueous solution

Classifying Solutions of Electrolytes

Electrolytes is any solutes which form ions when dissolved in water (or certain other solvents)

And it can be classified as strong and weak electrolyte.

1-*Strong electrolytes* ionize completely in a solvent and thus produce solutions that conduct electricity

2- ***Weak electrolytes*** ionize only partially. The solution of a weak electrolyte will not conduct electricity as well as a solution containing an equal concentration of a strong electrolyte.

Table 1 :Present the classification of strong and weak electrolytes

Classification of Electrolytes	
Strong	**Weak**
1. Inorganic acids such as HNO_3, $HClO_4$, H_2SO_4*, HCl, HI, HBr, $HClO_3$, $HBrO_3$ 2. Alkali and alkaline-earth hydroxides 3. Most salts	1. Many inorganic acids, including H_2CO_3, H_3BO_3, H_3PO_4, H_2S, H_2SO_3 2. Most organic acids 3. Ammonia and most organic bases 4. Halides, cyanides, and thiocyanates of Hg, Zn, and Cd

*H_2SO_4 is completely dissociated into HSO_4^- and H_3O^+ ions and for this reason is classified as a strong electrolyte. Note, however, that the HSO_4^- ion is a weak electrolyte and is only partially dissociated into SO_4^{2-} and H_3O^+.

Acid and Bases

Two definitions of acid and basis can be consider in analytical chemistry

1- Arrhenius definition
2- Bronsted-Lowry definition

Arrhenius definition

Arrhenius defined an acid as a substance that yields **hydrogen ions (H^+) when** dissolved **in water**. The hydrogen ion is a bare proton, but in water solution it is hydrated and exists as hydronium ion (H_3O^+).

$$H^{+1} + H_2O \rightarrow (H_3O)^{+1}$$

In dissolution process the highly polar water molecules react with the molecules of acid to form hydronium ions and the corresponding hydrated anion from the acid. This process is called ionization and refers to the formation of ions from atoms or molecules by the transfer of electrons.

Example:

$$HCl + H_2O \longrightarrow H_3O^{+1} + Cl^{-1}$$

$$H_2SO_4 + H_2O \longrightarrow H_3O^{+1} + HSO_4^{-1}$$

$$HNO_3 + H_2O \longrightarrow H_3O^{+1} + NO_3^{-1}$$

For base Arrhenius proposed that a base is a substance that yields **hydroxide ion (OH^{-1})** when dissolved in water.

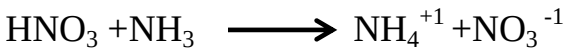
Example: Sodium hydroxide NaOH and Calcium hydroxide $Ca(OH)_2$

$$NaOH \longrightarrow Na^{+1} + OH^{-1}$$

Bronsted-Lowry definition

The Bronsted-Lowry concept defines an acid as any substance that can **give or donate a proton (H^{+1})** to some other substance.

A base is defined as any substance capable of **receiving or accepting a proton from some other substance.** In simple term, an acid is a proton donor, where a base is proton acceptor.



$$HNO_3 + NH_3 \longrightarrow NH_4^{+1} + NO_3^{-1}$$

Note:

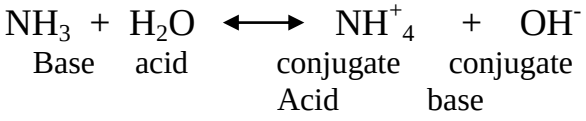
- A Bronsted-Lowry acid, like an Arrhenius acid, is a compound that breaks down to give an H+ in solution. The only difference is that the solution does not have to be water. We can still refer to the exact same acids as listed for the Arrhenius acid examples, but this time we'll change the solvent to ammonia, alcohol, or anything else.

- According to the Bronsted- Lowry theory, a base is a proton acceptor. According to the Arrhenius theory, a base is a hydroxide ion donor.

Conjugate Acids and Bases

*A **conjugate base*** is formed when an acid loses a proton. For example, acetate ion is the conjugate base of acetic acid. Similarly, ammonium ion is the conjugate acid of the base ammonia.

*A **conjugate acid*** is formed when a base accepts a proton.



Amphiprotic Species

Species that have both acidic and basic properties are **amphiprotic**. An example is dihydrogen phosphate ion, $H_2PO^-_4$, which behaves as a base in the presence of a proton donor such as H_3O^+

$$\underset{\text{Base}}{H_2PO^-_4} + \underset{\text{acid}}{H_3O^+} \longleftrightarrow \underset{\text{conjugate Acid}}{H_3PO_4} + \underset{\text{conjugate base}}{H_2O}$$

In the presence of a proton acceptor, such as hydroxide ion, however, $H_2PO^-_4$ behaves as an acid and donates a proton to form the conjugate base HPO_4

$$\underset{\text{acid}}{H_2PO^-_4} + \underset{\text{base}}{OH^-} \longleftrightarrow \underset{\text{conjugate base}}{HPO_4} + \underset{\text{conjugate acid}}{H_2O}$$

Water is the classic example of an **amphiprotic solvent**, that is, a solvent that can act either as an acid or as a base, depending on the solute. Other common amphiprotic solvents are methanol, ethanol AND asacetic acid.

$$\underset{\text{Base}}{NH_3} + \underset{\text{acid}}{CH_3OH} \longleftrightarrow \underset{\text{conjugate Acid}}{NH_4^+} + \underset{\text{conjugate base}}{CH_3O^-}$$

$$\underset{\text{Base}}{CH_3OH} + \underset{\text{acid}}{HNO_2} \longleftrightarrow \underset{\text{conjugate Acid}}{CH_3OH^+_2} + \underset{\text{conjugate base}}{NO_2^-}$$

$$HCl + CH_3C(O)(OH) \rightleftharpoons Cl^- + CH_3C(OH)_2^+$$

TABLE OF STRONG ACIDS

Completely Ionized in Water to Give One (*or more*) Protons per Acid Molecule

HI	$H^+(aq) + I^-(aq)$
HBr	$H^+(aq) + Br^-(aq)$
$HClO_4$	$H^+(aq) + ClO_4^-(aq)$
HCl	$H^+(aq) + Cl^-(aq)$
$HClO_3$	$H^+(aq) + ClO_3^-(aq)$
H_2SO_4	$H^+(aq) + HSO_4^-(aq)$ *(HSO_4^- is a weak acid that contributes additional protons)*
HNO_3	$H^+(aq) + NO_3^-(aq)$

TABLE OF STRONG BASES

Completely Ionized in Water to Give One (*or more*) Hydroxides per Base Molecule

NaOH	**$Na^+(aq) + OH^-(aq)$**
KOH	**$K^+(aq) + OH^-(aq)$**
LiOH	$Li^+(aq) + OH^-(aq)$
RbOH	$Rb^+(aq) + OH^-(aq)$
CsOH	$Cs^+(aq) + OH^-(aq)$
$Ca(OH)_2$	$Ca^{2+}(aq) + 2OH^-(aq)$ *(but not very soluble)*
$Ba(OH)_2$	$Ba^{2+}(aq) + 2OH^-(aq)$ *(but not very soluble)*
$Sr(OH)_2$	$Sr^{2+}(aq) + 2OH^-(aq)$ *(but not very soluble)*

EXERCISES

1. Label each reactant as a Brønsted-Lowry acid or a Brønsted-Lowry base.

$HCl(aq) + NH_3(aq) \rightarrow NH_4^+(aq) + Cl^-(aq)$

2. Label each reactant as a Brønsted-Lowry acid or a Brønsted-Lowry base.

$H_2O(\ell) + N_2H_4(aq) \rightarrow N_2H_5^+(aq) + OH^-(aq)$

3. Explain why a Brønsted-Lowry acid can be called a proton donor.
4. Explain why a Brønsted-Lowry base can be called a proton acceptor.
5. Write the chemical equation of the reaction of ammonia in water and label the Brønsted-Lowry acid and base.
6. Write the chemical equation of the reaction of methylamine (CH_3NH_2) in water and label the Brønsted-Lowry acid and base.
7. Demonstrate that the dissolution of HNO_3 in water is actually a Brønsted-Lowry acid-base reaction by describing it with a chemical equation and labeling the Brønsted-Lowry acid and base.
8. Identify the Brønsted-Lowry acid and base in the following chemical equation:

$C_3H_7NH_2(aq) + H_3O^+(aq) \rightarrow C_3H_7NH_3^+(aq) + H_2O(\ell)$

9. Write the chemical equation for the reaction that occurs when cocaine hydrochloride ($C_{17}H_{22}ClNO_4$) dissolves in water and donates a proton to a water molecule. (When hydrochlorides dissolve in water, they separate into chloride ions and the appropriate cation.)
10. If codeine hydrobromide has the formula $C_{18}H_{22}BrNO_3$, what is the formula of the parent compound codeine?

ANSWERS

1. HCl: Brønsted-Lowry acid; NH_3: Brønsted-Lowry base
2.
3. A Brønsted-Lowry acid gives away an H^+ ion—nominally, a proton—in an acid-base reaction.
4.
5. $NH_3 + H_2O \rightarrow NH_4^+ + OH^-$; NH_3: Brønsted-Lowry base; H_2O: Brønsted-Lowry acid
6.

7. $HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$; HNO_3: Brønsted-Lowry acid; H_2O: Brønsted-Lowry base
8.
9. $C_{17}H_{22}NO_{4+} + H_2O \rightarrow H_3O^+ + C_{17}H_{21}NO_4$

pH and pOH

-Adding an acid to water increases the H^+ ion concentration and decreases the OH^- ion concentration. Adding a base does the opposite.

-The product of the concentrations of these ions at equilibrium is always 1.0×10^{-14} at 25^oC.

$$[H^+][OH^-] = 1.0 \times 10^{-14}$$

-Since the concentrations of the H^+ and OH^- ions in aqueous solutions are usually smaller than 1 *M*, the logarithms of these concentrations are negative numbers

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

Similarly, **pOH** is the negative of the logarithm of the OH^- ion concentration.

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

-The sum of pH and pOH for any solution will equal 14

$$pH + pOH = 14$$



Acidic Neutral Basic

What is the pH of a solution whose $[H^+] = 1 \text{ X } 10^{-4}$ M

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

$pH = - \log[1 \text{ X } 10^{-4}]$

$pH = - [\log 1 + \log 10^{-4}]$

Note: When you multiply numbers you always ADD their log forms

log 1 is always zero

$\log 10^{x} = x$ so $\log 10^{-4} = -4$

$pH = - [\log 1 + \log 10^{-4}] = - [0 + (-4)] = - [-4] = \mathbf{+4}$

Calculate the pH of a solution that has a $[OH^{1-}] = 1 \text{ X } 10^{-5}$ M

Determine $pOH = -\log[OH^{1-}] = -\log [1 \text{ X } 10^{-5}] = 5$

Use the relationship pH + pOH = 14

pH + 5 = 14

pH = 14 -5 = 9

What Is Chemical Equilibrium?

Chemical equilibrium refers to the state of a system in which the concentration of the reactant and the concentration of the products do not change with time, and the system does not display any further change in properties.

When the rate of the forward reaction is equal to the rate of the reverse reaction, the state of chemical equilibrium is achieved by the system. When there is no further change in the concentrations of the reactants and the products due to the equal rates of the forward and reverse reactions, the system is said to be in a state of dynamic equilibrium.

Types of Chemical Equilibrium

1-Precipitation Reactions

2-Acid–Base Reactions

3-Oxidation–Reduction (Redox) Reactions

4-Complex reactions

Furthermore chemical equilibrium can be divided according to its phase i.e

- **Homogeneous Equilibrium**
- **Heterogeneous Equilibrium**

Homogenous Chemical Equilibrium

In this type, the reactants and the products of chemical equilibrium are all in the same phase. Homogenous equilibrium can be further divided into two types:

Reactions in which the number of molecules of the products is equal to the number of molecules of the reactants. For example,

- **H_2 (g) + I_2 (g) ⇌ 2HI (g)**
- **N_2 (g) + O_2 (g) ⇌ 2NO (g)**

Reactions in which the number of molecules of the products is not equal to the total number of reactant molecules. For example,

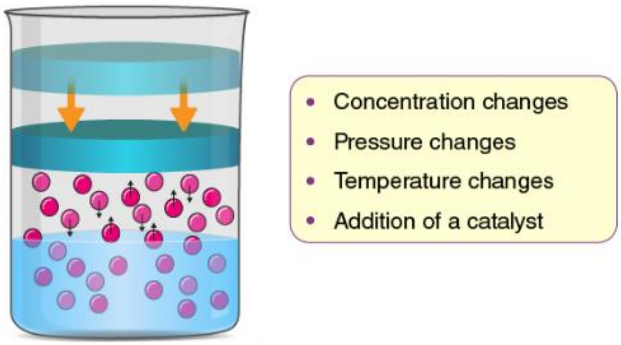
- **$2SO_2$ (g) + O_2 (g) ⇌ $2SO_3$ (g)**
- **$COCl_2$ (g) ⇌ CO (g) + Cl_2 (g)**

Heterogeneous Chemical Equilibrium

In this type, the reactants and the products of chemical equilibrium are present in different phases. A few examples of heterogeneous equilibrium are given below:

- **CO_2 (g) + C (s) ⇌ 2CO (g)**
- **$CaCO_3$ (s) ⇌ CaO (s) + CO_2 (g)**

Thus, the different types of chemical equilibrium are based on the phase of the reactants and products.

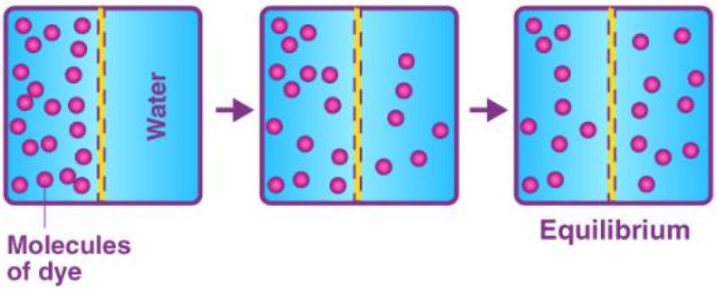


Factors Affecting Equilibrium

Chemical Equilibrium – Factors Affecting Equilibrium

Le-Chatelier's principle

According to **Le-Chatelier's principle**, if there is any change in the factors affecting the equilibrium conditions, the system will counteract or reduce the effect of the overall transformation. This principle applies to both chemical and **physical equilibrium** (It represents the existence of the same substance in two different physical states, in this process, there is no change in chemical composition. such as the solubility of solid dye in water).



Factors Affecting Equilibrium

There are several factors like temperature, pressure and concentration of the system that affect equilibrium. Some important factors affecting chemical equilibrium are :

Change in Concentration

As per Le Chatelier's principles, the only way for equilibrium to accept more reactants is to increase product formation. The forward reaction is favored when the concentration of the reactant is increased. The equilibrium of the reaction shift towards the use of reactants in the reaction, which decreases the concentration of the reactants.

Similarly, the addition of product (concentration/pressure) shall increase the backward reaction to decrease the product concentration. The backward reaction is favoured when the concentration of the reactant decreases, the equilibrium of the reaction shift towards the production of reactants, and the concentration of the reactants is more.

Example:

Consider a reaction between oxygen and sulfur dioxide to produce sulfur trioxide.

$2SO_{2(g)} + O_{2(g)} \leftrightharpoons 2SO_{3(g)}$

If the concentration of the reactant increases, then

- Equilibrium will shift towards the decrease in the concentration of the reactants.
- More favored for the forward reaction.
- Some of the SO_2 or O_2 form SO_3.
- Equilibrium of the reaction shift towards the right.

If the concentration of the reactant decreases, then

- Equilibrium will shift towards the increase in the concentration of the reactants.
- More favored for the backward reaction.
- Some of the SO_3 would change to SO_2 or O_2.
- Equilibrium of the reaction shift towards the left.

If the product decreases, then

- The equilibrium of the reaction shift to increase the concentration of the sulfur trioxide.
- Increase in the forward reaction rate.
- Some of the SO_2 or O_2 form SO_3.
- Equilibrium of the reaction shift towards the right.

If the product increases, then

- The equilibrium of the reaction shifts to decrease the concentration of the sulfur trioxide.
- Increase in the reverse reaction rate.
- Some of the SO_3 would change to SO_2 or O_2.
- Equilibrium of the reaction shift towards the left.

Change in Pressure

Change in pressure happens due to the change in the volume. If there is a change in pressure, it can affect the gaseous reaction as the total number of gaseous reactants and products are now different. According to Le Chatelier's principle, in heterogeneous chemical equilibrium, the change of pressure in both liquids and solids can be ignored because the volume is independent of pressure.

Change in Temperature

The effect of temperature on chemical equilibrium depends upon the sign of ΔH of the reaction and follows Le-Chatelier's Principle.

Along with the equilibrium constant, the rate of reaction is also affected by the change in temperature. **As per Le Chatelier's principle, the equilibrium shifts towards the reactant side when the temperature increases in the case of exothermic reactions; for endothermic reactions, the equilibrium shifts towards the product side with an increase in temperature.**

The individual reaction in the equilibrium can be either endothermic or exothermic. Likewise, at equilibrium, the net energy involved may make the reversible reactions either endothermic or exothermic.

According to Le Chatelier's Principles,

- In exothermic equilibrium, an increase in temperature decreases product formation, and a decrease in temperature increases product formation.
- In endothermic reactions, an increase in temperature increases the product formation, and a decrease in temperature decreases the product formation

$$\log\left[\frac{K_1}{K_2}\right] = \frac{-\Delta H^\circ}{R}\left[\frac{1}{T_2} - \frac{1}{T_1}\right]$$

Le Chatelier's Principle on Change of Temperature

As per the Van't Hoff equation, for an exothermic equilibrium, ΔH will be negative. An increase in temperature shall decrease K_2, or a decrease in temperature increases K_2. The opposite is true for an endothermic reaction.

Example:

Consider a reaction

$N_{2(g)} + 3H_{2(g)} \leftrightharpoons 2NH_{3(g)}$ $\Delta H = -92kJ$

Increase in temperature

- This prefers an endothermic reaction because it takes energy.
- An endothermic reaction is a reverse reaction, and it is favored.
- The yield of the product (NH_3) decreases.

Decrease in temperature

- This prefers an exothermic reaction because it gives energy.
- An exothermic reaction is a forward reaction, and it is favored.
- The yield of the product (NH_3) increases.

Effect of a Catalyst

A catalyst does not affect the chemical equilibrium, it only speeds up a reaction. In fact, the catalyst equally speeds up the forward as well as the reverse reaction. This results in the reaction reaching its equilibrium faster.

The same amount of reactants and products will be present at equilibrium in a catalysed or non-catalysed reaction. The presence of a catalyst only facilitates the reaction to proceed through a lower-energy transition state of reactants to products.

Effect of Addition of an Inert Gas

When an inert gas like argon is added to a constant volume, it does not take part in the reaction, so the equilibrium remains undisturbed. If the gas added is a reactant or product involved in the reaction, then the reaction quotient will change.

Acid and base equilibrium constant

For acid base reactions, the equilibrium constant can be divided into acid dissociation constant, K_a and base dissociation constant K_b

The reaction of an acid with its solvent (typically water) is a strong acid dissociation reaction.

$HCl_{(aq)} + H_2O_{(l)} \longrightarrow H_3O_{(aq)} + Cl^-_{(l)}$

Use a single arrow $\longrightarrow$ in place of the equilibrium arrow (↔) because we treat HCl as if it dissociates completely in an aqueous solution.

The equilibrium constant for this reaction is an acid dissociation constant, *K*a, which we write as

$$K_a = \frac{[Cl-][H3O+]}{[HCl]}$$

A weak acid, of which aqueous acetic acid is one example, does not completely donate its acidic proton to the solvent. Instead, most of the acid remains undissociated with only a small fraction present as the conjugate base.

$CH_3COOH_{(aq)} + H_2O_{(l)} \longleftrightarrow H_3O^+_{(aq)} + CH_3COO^-_{(aq)}$

The equilibrium constant for this reaction is an acid dissociation constant, *K*a, which we write as

$$K_a = \frac{[CH3COO-][H3O+]}{[CH3COOH]}$$

- A **monoprotic** weak acid, such as acetic acid, has only a single acidic proton and a single acid dissociation constant. Other acids, such as phosphoric acid, have multiple acidic protons, each characterized by an acid dissociation constant. We call such acids **polyprotic**. Phosphoric acid, for example, has three acid dissociation reactions and three acid dissociation constants.

$H_3PO_{4\,(aq)} + H_2O_{(l)} \longleftrightarrow H_3O^+_{(aq)} + H_2PO_4^-{}_{(aq)}$

$$K_{a1} = \frac{[H2PO4-][H3O+]}{[H3PO4]} = 7.11 * 10^{-3}$$

$H_2PO_4^-{}_{(aq)} + H_2O_{(l)} \longleftrightarrow H_3O^+_{(aq)} + HPO_4^{-2}{}_{(aq)}$

$$K_{a2} = \frac{[HPO4\text{^}{-2}][H3O+]}{[H2PO4-]} = 6.32 * 10^{-8}$$

$$HPO_4^{-2}{}_{(aq)} + H_2O_{(l)} \longleftrightarrow H_3O^+_{(aq)} + PO_4^{-3}{}_{(aq)}$$

$$K_{a3} = \frac{[PO4\text{^}{-3}][H3O+]}{[HPO4\text{^}{-2}]} = 4.5 * 10^{-13}$$

The decrease in the acid dissociation constants from *K*a1 to *K*a3 tells us that each successive proton is harder to remove. Consequently, H_3PO_4 is a stronger acid than $H_2PO_4^-$, and $H_2PO_4^-$ is a stronger acid than HPO_4^{2-} .

- The most common example of a strong base is an alkali metal hydroxide, such as sodium hydroxide, NaOH, which completely dissociates to produce hydroxide ion.

$$NaOH_{(s)} \longrightarrow Na^+_{(aq)} + OH^-_{(aq)}$$

$$k_b = [Na^+][OH^-]$$

1. Write equilibrium constant expressions for the following reactions. What is the value for each reaction's equilibrium constant?

   a. $NH_3(aq) + H_3O^+(aq) = NH_4^+(aq)$

   b. $PbI_2(s) + S^{2-}(aq) = PbS(s) + 2I^-(aq)$

   c $CdY^{2-}(aq) + 4CN^-(aq) = Cd(CN)_4^{2-}(aq) + Y^{4-}(aq)$

   d. $AgCl(s) + 2NH_3(aq) = Ag(NH_3)_2^+(aq) + Cl^-(aq)$

   e. $BaCO_3(s) + 2H_3O^+(aq) = Ba^{2+}(aq) + H_2CO_3(aq) + 2H_2O(l)$

Expressing Concentrations of Solutions

A. Concentration percent by Mass

$$\text{Percent by mass} = \frac{\text{mass of solute}}{\text{mass of solution}} \text{ X } 100\%$$

Mass of solution = Mass of solute + Mass of solvent

Example 1: Calculate the percent of NaCl if 19 g of NaCl is dissolved in enough water to make 175 g of solution.

Solution:

$$\text{Percent by mass} = \frac{\text{mass of solute}}{\text{mass of solution}} \text{ X } 100\%$$

$$\textbf{Percent by mass} = \frac{\mathbf{19}}{\mathbf{175}} \times \mathbf{100}$$

$$= \mathbf{10.85\%}$$

Example:

How many grams of glucose and of water are in 500 g of a 5.3% by-mass glucose solution

Solution

5.3 % Means 5.3 g of solute/100 g of solution

IF 5.3 g of glucose needs 100 g glucose solution

THEN (x) needs of glucose in 500 g of glucose solution

X= 26.5 g of glucose

Mass of solvent = 500 – 26.5

= 437.5 g

Example 3: Calculate the number of grams of sugar that must be dissolved in 825 g of water to prepare 20% sugar solution.

Solution

20 % Means 20 g of solute/100 g of solution

Mass of solution = Mass of solute + Mass of solvent

Mass of solvent = Mass of solution - Mass of solute

= 100 – 20

= 80 g of H2O

IF 20 g of SUGAR needs 80 g of H2O

THEN (x) needs of SUGAR in 825 g of H2O

X = 206 g of SUGAR

C. Concentration in Parts per Million (ppm) and parts per Billion (ppb =1microgram/L)

$$\text{Part per million (ppm)} = \frac{\text{mass of solute}}{\text{mass of solution}} \text{ X } 1{,}000{,}000$$

Example: A water contain 3.5 mg of fluride (F^{-1}) ions in 825 ml of the water sample. Calculate the ppm of the fluride ion in the sample. (Assume that the density of the very dilute water sample is 1.0 g/ml).

Solution:

$$\text{Part per million (ppm)} = \frac{3.5 \text{ g}}{825 \text{ ml}} \text{ X } \frac{1\ ml}{1\ g\ sample} \times \frac{1\ g\ sample}{1000\ mg\ sample} \times 1{,}000{,}000$$

= 4.2 ppm

D. Concentration in Molarity

$$\text{Molarity (M)} = \frac{\text{moles solute}}{\text{volume (liter) solution}} = \frac{\text{millimoles solute}}{\text{milliliter solution}}$$

Example #1: Calculate the molarity of a H_3PO_4 solution containing 284 g of H3PO4 in 2.2 L of solution.

M.wt of H_3PO_4 = 98

$$\text{Molarity (M)} = \frac{\text{moles solute}}{\text{volume (liter) solution}}$$

Moles of solute = $\frac{Wt}{Mwt} = \frac{284\ g}{98\ g/mol} = 2.89$ mol

M = $\frac{2.89\ mol}{2.2\ L} = 1.32$ M

Example #2:

How many moles of hydrochloric acid are in 200 mL of 0.15 M HCl?

Solution:

$$\text{Molarity (M)} = \frac{\text{moles solute}}{\text{volume (liter) solution}}$$

$$0.15 = \frac{\text{moles solute}}{1\ \text{L}}$$

Moles of solute = 0.15 mol

IF Moles of HCl = 0.15 in 1 L (1000 ml)

THEN (X) moles in 200 ml = X moles of HCl is 0.03 M

Example#3: What mass of sodium hydroxide (NaOH) is needed to prepare 100 ML of 0.125 M sodium hydroxide?

Solution:

Mwt NaOH = 40

IF Moles of NaOH 0.125 mol in 1000 ml

THEN Moles of NaOH (X) in 100 ml WILL BE = 0.0125 mol

Moles of solute = $\frac{Wt}{Mwt}$

= 40 g/mol * 0.0125 mol = 0.5 g

E. Concentration by Molality

m = molality= $\frac{moles\ of\ solute}{kilograms\ of\ solvent}$

Example #1: Calculate the molality of H_3PO_4 solution containing 32.7 g of H_3PO_4 in 100 g of water.

Solution:

m = molality= $\frac{moles\ of\ solute}{kilograms\ of\ solvent}$

Mwt of H_3PO_4 = 98 g/mol

Moles of H3PO4 = $\frac{Wt}{Mwt}$

= $\frac{32.7\ g}{98\ g/mol}$ = 0.33 mol

molality = $\frac{0.33}{100\ g} \times \frac{1000\ g}{1Kg}$

= 3.34 m

Example #2: Calculate the number of grams of glycerol ($C_3H_8O_3$) necessary to prepare 520 g of 1 molal solution of glycerol in water.

Solution:

Wt of $C_3H_8O_3$= 92

m = molality= $\frac{moles\ of\ solute}{kilograms\ of\ solvent}$

No. of moles of $C_3H_8O_3$= 1 mol (92g)

Total mass of glycerol solution = 1000 +92

=1092 g

IF 1092 g of glycerol solution gives 92 g glycerol

THEN 520 g of glycerol solution gives (X) g of glycerol

X= 43.8 g

F. Concentration by Normality

N= Normality = $\frac{equivalent\ of\ solute}{liter\ of\ solution}$

A gram equivalent weight or equivalent is a measure of the reactive capacity of a given chemical species (ion, molecule, etc.). Normality is the only concentration unit that is reaction dependent. **This type of concentration expression widely used in reactions of acid-base or oxidation-reduction reactions.**

One equivalent of any acid is = $\frac{grams-formula\ mass\ of\ acid}{no.of\ H\ ions\ capable\ of\ supplying\ by\ acid}$

Example

H_2SO_4= $\frac{grams-formula\ mass\ of\ acid}{no.of\ H\ ions\ capable\ of\ supplying\ by\ acid} = \frac{98}{2} = 49$

$Ca(OH)_2$ if 1 OH^- is replaced= $\frac{grams-formula\ mass\ of\ base}{no.of\ OH\ ions\ capable\ of\ supplying\ by\ base} = \frac{74}{1} = 74$

Example 1 : Calculate the normality of H_3PO_4 solution containing 275 g of H_3PO_4 in 1.2 L of solution in reactions that replace all three hydrogen ions.

Solution:

N= Normality = $\frac{equivalent\ of\ solute}{liter\ of\ solution}$

The grams-formula mass of H_3PO_4 is 98 g.

Since 3moles of H^+ are used per mol of acid then

One equivalent of H_3PO_4 = $\frac{98}{3}$ = 32.7 g

$$\frac{275\ g\ H3PO4}{1.2\ L} \times \frac{1\ equivalent\ H3PO4}{32.7\ g\ H3PO4} = \frac{7.01\ eq}{1\ L\ solution} = 7.01\ N$$

Example 2: Calculate the number of grams of H2SO4 necessary to prepare 520 ml of 0.1N sulfuric acid in reaction that replace both hydrogen ions.

The grams-formula mass of H_2SO_4 is 98.1 g.

Since 2moles of H+ are used per mol of acid then

One equivalent of H_3PO_4 = $\frac{98.1}{2}$ = 49 g

N= Normality = $\frac{equivalent\ of\ solute}{liter\ of\ solution}$

0.1$\frac{equivalent\ of\ solute}{1\ L\ of\ solution}$* 520 ml * $\frac{1\ L}{1000ml}$= 0.052 equivalent of solute

0.052 equivalent of solute *$\frac{49\ g\ of\ H2SO4}{1\ equivalent\ of\ H2SO4}$ = 2.55 g of H2SO4

Example: Calculate the percent by mass of the following aqueous solution

a- 6.8 *m* sulfuric acid solution

$$\text{Percent by mass} = \frac{\text{mass of solute}}{\text{mass of solution}} \text{ X } 100\%$$

Mass of solute = 6.8 moles of H_2SO_4 * $\frac{98.1\ g\ H2SO4}{1\ mol\ H2SO4}$

Mass of solution = 667 g H_2SO_4+ 1000 g H_2O

= 1667 g Solution

$$\text{Percent by mass} = \frac{667 \text{ g H2SO4}}{1667 \text{ g solution}} \text{ X } 100\%$$

Structure of atom

atom

NUCLEUS

The nucleus is the **center of mass** but does not significantly contribute to volume.

It is made up of:

Protons: Mass = 1 amu, charge = +1

Neutrons: Mass = 1 amu, charge = 0

ELECTRONS

The electronic cloud determines the size, or volume, of the atom, but does not significantly contribute to mass

ELECTRONS: Mass = .0005 amu (negligible), charge = -1

Electrons are found outside the nucleus

They occupy orbitals

They are the units of negative charge

They determine the atomic number (Z)

The atomic number determines the atom’s identity

General arrangement of electrons, protons and neutrons:

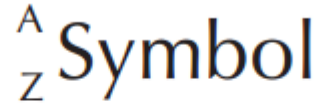
1- All protons and neutrons are found in nucleus, the protons are positively charged and neutrons are neutral, thus the relative charge of nucleus must be positive and equal to the No. of protons.
2- An atom is electrically neutral, hence the number of protons = the number of electron. If the number of proton dose not equal the number of electron the atom is positively or negatively charged.
3- Electrons are found outside the nucleus in shells of a certain energy level. The electron in an atom have their energy controlled to certain energy values called (Energy Levels). For an electron to change its energy, it must be shift from one energy level to another, thus an amount of energy is required equal to the energy difference between the two levels. Keep in mined that the electrons can not go to lower energy levels because they are arranged in their lowest energy level.

Mass and atomic number

To the left of the symbol we find the atomic mass (A) at the upper corner, and the atomic number (Z) at the lower corner.

The **mass number** (**A**), also called **atomic mass number** or **nucleon number**, is the total number of protons and neutrons (together known as nucleons) in an atomic nucleus.

Atomic number is the number of protons found in the nucleus of an atom of that element, and therefore identical to the charge number of the nucleus. It is conventionally represented by the symbol ***Z***.



Examples: $^{1}_{1}H$ $^{12}_{6}C$ $^{23}_{11}Na$

- **No two different elements can have the same atomic number.** The periodic table is arranged by order of increasing atomic number, which is always an integer.

Atomic number = number of protons in nucleus= number of electron

- **Different forms of the same element can have different masses. They are called isotopes**. The following are representations for some of the isotopes of hydrogen and carbon.

$^{1}_{1}H$	$^{2}_{1}H$	$^{12}_{6}C$	$^{13}_{6}C$
regular hydrogen	deuterium	regular carbon, or C-12	carbon-13, or C-13

Mass number = sum of the protons + neutrons

Examples:

$^{4}_{2}\mathbf{He}$

2= atomic number= number of protons in nucleus

4= mass number= sum of the protons + neutrons

4-2=2 number of neutrons in nucleus

Since mass of elec. Is negligible in terms of atomic mass

Number of electron = number of protons =2 electron outside the nucleus

a- $^{27}_{13}Al$

13= atomic number= number of protons in nucleus

27= mass number= sum of the protons + neutrons

27-13=14 number of neutrons in nucleus

Since mass of elec. Is negligible in terms of atomic mass

Number of electron = number of protons =13 electron outside the nucleus

$^{268}_{105}$Db (Dubnium)

105= atomic number= number of protons in nucleus

268= mass number= sum of the protons + neutrons

268-105=163 number of neutrons in nucleus

Number of electron = number of protons =105 electron outside the nucleus

Arrangement of electrons in principal energy level

The total number of electrons in the atom can be obtained from the atomic number.

The electrons can exist at principal energy levels and are arranged in shells, which increase in energy as they increase in distance from the nucleus.

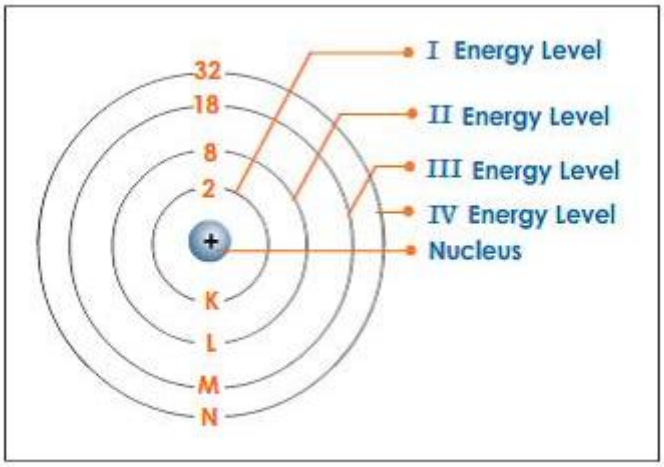
Electrons in atom occupy the lowest energy that are available in the beginning. The nearer the electron is to the nucleus the less energy the electron has.

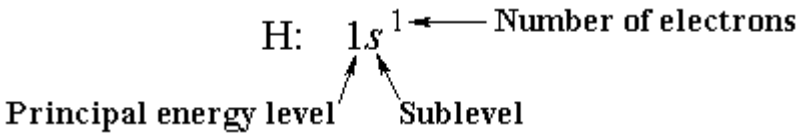
The principal energy levels are express by numbers 1,2,3,4,5,6 and 7.

Maximum numbers of electrons in a given energy level is found from the following equation:

Maximum number of electron in principal energy levels = $2n^2$

n = integers 1 to 7 of the principal energy levels





Examples:

1- Calculate the maximum number of electrons at the 1 and 2 principal energy level

Maximum No. of electrons = $2\times1^2 = 2$

Maximum No. of electrons = $2\times2^2 = 8$

2- Calculate the maximum number of electrons in helium atom and write the electronic configuration

$\mathbf{^4_2He}$

2= atomic number= number of protons in nucleus

Number of electron = number of protons =2 electron outside the nucleus

Electronic configuration = $1s^2$ principal energy level 1 filled with (2 valance electron)

$\mathbf{^7_3Li}$ = $1s^2\ 2s^1$ (1 valance electron)

$\mathbf{^{11}_5B}$ = $1S^2\ 2S^2\ 2P^1$ (3 valance electron)

$\mathbf{^{39}_{19}K}$ = $1S^2, 2S^2\ 2P^6, 3S^2 3P^6, 4S^1$ (1 valance electron)

Ionic charge for single atoms

In neutral atom:

Number of proton (positive charge) = Number of electron (negative charge)

However, when the atom often gain or lose electrons, this atom has no longer the same number of protons and electrons thus it called **charged atom**

Charged atom can hold positive or negative charge, ions with negative charge are called **anions** (are formed when a neutral atom gain electrons). Remember gaining electrons is gaining negative charge.

Ions with a positive charge are called **cations**. Cations are formed when a neutral atom loses electrons.

Charge= No. of protons – No. of e^-

Problems:

Calculate the charge on the following ion and indicate whether the ions are cations or anions

Ion	Protons	Electrons
X	**35**	**36**
Y	**3**	**2**
Z	**20**	**18**

Solution

X Charge= +35-36 = -1 (anion) **Y Charge= +3 -2 = +1(cation)**

Z Charge= +20-18 = +2(cation)

Problem 2

Calculate the number of protons, neutrons and electrons in the following ions

a- $^{24}_{12}Mg^{+2}$

No. of P = atomic no.= 12

No. of N = mass no. - atomic no. = 12

Charge = No. of P – No. of e^- ⟶ Charge – No. of P = - No. of e^-

+2 – 12 = -10 electrons (in Mg^{+2})

b- $^{14}_{7}N^{-3}$

No. of P = atomic no.= 7

No. of N = mass no. - atomic no. = 7

Charge = No. of P – No. of e- ⟶ Charge – No. of P = - No. of e-

-3 – 7 = -10 electrons (in N^{-3})

Titration

What is titration?

Titration is an analytical method primarily used in determining the concentration of an unknown sample. Titration is a procedure for determining the concentration of an acid or base in solution through the addition of base or an acid of known concentration until the neutralization point **(End Point)** is reached as shown by an indicator.

Indicators often used in titrations such as methyl orange, methyl red, phenolphthlein.

The neutralization can be represented by a general equation,

$HX + BOH \longrightarrow BX + HOH$

Where HX is an acid and BOH is a base. Water is one of the products.

Titrations can usually occur in reactions such as **acid-base reactions.** During the process, two important stages known as **endpoint and equivalence point** are reached. A point of **equivalence** in a titration refers to a point at which the added titrant is chemically equivalent to the sample analyte. On the other side, **endpoint** is a point where the solution changes colour. The main difference between equivalence and endpoint is that the equivalence point is a point where the chemical reaction comes to an end, while the endpoint is the point where the colour change occurs in a system.

The procedure of titration

In titration, a measured amount of the acid or base of unknown concentration is placed in a flask and a drop or two of an indicator is added. To this solution, a solution of acid or base of known concentration is slowly added from a burette until the color of the indicator just changed. This called the end point.

Note: The known solution is called the titrant and the unknown solution is called the analyte.

When the mole ratio =1 then you can use $M_1V_1=M_2V_2$

If the mole ratio not equal to1 then you have to solve the problem according to mole ratio

Problem 1:

A 25.0 mL sample of an unknown acid is titrated with 0.100 M NaOH. It takes 40.0 mL of the NaOH solution to neutralize the acid. What is the concentration of the unknown acid?

Solution:

We can use the following equation to calculate the concentration of the unknown acid:

M1V1 = M2V2

where:

M1 is the concentration of the known base (NaOH)

V1 is the volume of the known base solution used (40.0 mL)

M2 is the concentration of the unknown acid

V2 is the volume of the unknown acid solution titrated (25.0 mL)

Substituting the known values into the equation, we get:

0.100 M * 40.0 mL = M2 * 25.0 mL

Solving for M2, we get:

M2 = 0.160 M

Therefore, the concentration of the unknown acid is 0.160 M.

Problem 2:

A solution of 0.500 M HCl is titrated with 0.250 M NaOH. If the initial volume of the HCl solution is 50.0 mL, what volume of the NaOH solution is needed to neutralize the acid?

Solution:

We can use the same equation from the previous problem to calculate the volume of the NaOH solution needed to neutralize the acid:

M1V1 = M2V2

where:

M1 is the concentration of the known acid (HCl)

V1 is the initial volume of the known acid solution (50.0 mL)

M2 is the concentration of the known base (NaOH)

V2 is the volume of the known base solution needed to neutralize the acid

Substituting the known values into the equation, we get:

0.500 M * 50.0 mL = 0.250 M * V2

Solving for V2, we get:

V2 = 100.0 mL

Problem3:

A 25.00 mL sample of an unknown acid solution is titrated with a 0.200 M NaOH solution. The endpoint is reached after 37.50 mL of the NaOH solution is added. What is the molarity of the unknown acid solution?

Solution:

First, write the balanced chemical equation for the reaction between the acid and the NaOH:

$HA + NaOH \rightarrow NaA + H_2O$

From the balanced chemical equation, we can see that the mol ratio between the acid (HA) and the NaOH is 1:1. This means that 1 mole of HA will react with 1 mole of NaOH.

Next, calculate the number of moles of NaOH used in the titration:

moles of NaOH = M × V

where:

- NaOH solution used (37.50 mL)

moles of NaOH = 0.200 M × 37.50 mL
moles of NaOH = 0.0750 mol

Since the mol ratio between the acid and the NaOH is 1:1, the number of moles of acid is equal to the number of moles of NaOH:

moles of acid = moles of NaOH = 0.0750 mol

Finally, calculate the molarity of the acid solution:

Molarity of acid = moles of acid / volume of acid solution

where:

- moles of acid = 0.0750 mol
- volume of acid solution = 25.00 mL

Molarity of acid = 0.0750 mol / 25.00 mL
Molarity of acid = 0.300 M

Therefore, the molarity of the unknown acid solution is 0.300 M.

Problem 4: IF it takes 50 ml of 0.5 Ml OF 0.5 M KOH solution to completely neutralize 125 ml of sulfuric acid solution. What is the concentration of H2SO4 solution.

$H_2SO_4 + 2KOH \longrightarrow K_2SO_4 + 2H2O$

From the balanced chemical equation, we can see that the mol ratio between the acid (H_2SO_4) and the KOH is 1:2. This means that 1 mole of KOH will react with 0.5 mole of H_2SO_4.

Next, BY calculation the number of moles of KOH used in the titration:

n(moles of KOH) = M × V = 0.5*0.05 = 0.025 mol

n(H_2SO_4)= ½(KOH)= 0.0125 MOL

where:

- moles of acid = 0.0125mol
- volume of acid solution = 0.125 L

Molarity of acid = 0.0125mol / 0.125L

Molarity of acid = 0.1 mol / L

Therefore, the molarity of the unknown acid solution is 0.100 M.

EX: A 10.0 mL sample of a mixture of acetic acid (CH3COOH) and hydrochloric acid (HCl) is titrated with 0.250 M NaOH solution. 30.0 mL of the NaOH solution is required to reach the equivalence point. What is the concentration of the acetic acid solution and the hydrochloric acid solution in the mixture?

To solve this titration problem, we will use the following steps:

1. Write the balanced chemical equations for the reactions between acetic acid and sodium hydroxide, and hydrochloric acid and sodium hydroxide.

$CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$

$HCl + NaOH \rightarrow NaCl + H_2O$

2. Determine the moles of sodium hydroxide used.

0.250 mol/L NaOH * 0.0300 L NaOH = 0.00750 mol NaOH

3. Use the balanced chemical equations to determine the moles of acetic acid and hydrochloric acid that reacted with the sodium hydroxide.

Since the sodium hydroxide solution is the titrant, we know that it reacted with all of the acetic acid and hydrochloric acid in the mixture. Therefore, the moles of acetic acid and hydrochloric acid that reacted with the sodium hydroxide are equal to the moles of sodium hydroxide used.

0.00750 mol NaOH * (1 mol CH_3COOH / 1 mol NaOH) = 0.00750 mol CH3COOH

0.00750 mol NaOH * (1 mol HCl / 1 mol NaOH) = 0.00750 mol HCl

4. Calculate the concentrations of the acetic acid solution and the hydrochloric acid solution in the mixture.

We can calculate the concentrations of the acetic acid solution and the hydrochloric acid solution in the mixture by dividing the moles of each acid by the volume of the mixture.

[CH_3COOH] = 0.00750 mol CH_3COOH / 0.0100 L = 0.750 M CH_3COOH

[HCl] = 0.00750 mol HCl / 0.0100 L = 0.750 M HCl

Therefore, the concentration of the acetic acid solution and the hydrochloric acid solution in the mixture is 0.750 M.

It is important to note that this problem assumes that the acetic acid and hydrochloric acid in the mixture react completely with the sodium hydroxide.

Problems:

1. A 50.0 mL sample of an unknown base is titrated with 0.200 M HCl solution. 25.0 mL of the HCl solution is required to reach the equivalence point. What is the concentration of the unknown base?
2. A 20.0 mL sample of phosphoric acid (H3PO4) is titrated with 0.150 M NaOH solution. 40.0 mL of the NaOH solution is required to reach the equivalence point. What is the concentration of the phosphoric acid solution?

Answers:

1. 0.100 M
2. 0.300 M