Chapter one

Introduction

Chemistry: The sience deals with the study of composition, structure and the properties of substance.

Science of chemistry can be classified into:

a) Organic chemistry.
b) Inorganic chemistry.
c) Biochemistry.
d) Physical chemistry.
e) Analytical chemistry.

Analytical chemistry: deals with methods of separation and identification of one or more of the component(s) in a sample and determination of the relative amounts of each.

Analytical chemistry is classified depends upon, weight of sample, volume of sample and purpose of analysis.

1) weight of sample: several of analytical methods are suitable for:

a) macro analysis: when the amount of sample (solid) is 0.05-0.1 gm.
b) semimicro analysis: when the amount of sample (solid) is 0.005-0.01 gm.
c) micro analysis: when the amount of sample (solid) is 0.0005-0.001 gm.
2) Purpose of analysis: includes
   a) Qualitative analysis.
   b) Quantitative analysis.
   c) Absorption of photo energy.
   d) Emission of photo energy.
   e) Analysis of gases.
   f) Electrochemical methods.
   g) Chromatographic method.

   **Analyte**: The component(s) of the sample that are to be determined.

Analytical chemistry may be qualitative or quantitative:

1) *Qualitative analysis*: identification process used to detect anions and cations in an analyte, without care to its amounts.

2) *Quantitative analysis*: determination of quantity of an analyte in measured amount of a sample. The results of quantitative analysis are expressed in terms as parts per hundred, per thousands and per million.

   Quantitative analysis can be done by:
   
   a) Volumetric methods.
   
   b) Gravimetric methods.

   **Volumetric analysis**: is the measure of volume of standard solution that will be equivalent the volume of unknown solution (analyte).

   There are four reactions occur in volumetric analysis:
   
   a) Acid-base titration.
b) Precipitation titration.

c) Oxidation-reduction titration.

d) Complex formation titration.

**Gravimetric analysis**: is the precipitate, separate and weight of analyte by using:

a) Organic precipitate.

b) Inorganic precipitate.

c) Electronic deposition.

**Steps of chemical analysis**:

a) Choose the method of analysis.

b) Sampling.

c) Preparation of sample for analysis.

d) Measurement of sample.

e) Dissolution of sample.

f) Separation of interferences.

g) Calculations and conclusions.

**Chemical equilibrium**:

There are two kinds of reactions:

1) Reversible reaction, \[ A + B \rightleftharpoons C + D \]

\[ \text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^- \]

2) Irreversible reaction, \[ A + B \rightarrow C + D \]

\[ \text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O} \]
**Reaction which proceed nearly to competition:**

1) When the reaction gives slightly soluble precipitate:

\[ \text{Pb}^{+2} + \text{S}^{-2} \rightarrow \text{PbS} \]

\[ \text{Cu}^{+2} + \text{S}^{-2} \rightarrow \text{CuS} \]

2) When one of the products is a gas (formation of gas):

\[ \text{CaCO}_3 + 2\text{HCl} \rightarrow \text{CO}_2 \uparrow + \text{CaCl}_2 + \text{H}_2\text{O} \]

\[ \text{FeS} + 2\text{H}^+ \rightarrow \text{H}_2\text{S} + \text{Fe}^{+2} \]

3) When the product is highly stable complex ion:

\[ \text{Hg}^{+2} + 4\text{I}^- \rightarrow \text{HgI}_4^{-2} \]

\[ \text{AgCl} + 2\text{NH}_3 \rightarrow (\text{AgNH}_3)_2^{+2} \]

4) When the reaction gives a very weak electrolite (or nonionized product):

\[ \text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O} \]

where

is (H\(_2\)O) is a very weak electrolite.

In generally, most of chemical reactions are consider as reversible reactions.

There are two kinds of system:

a) Homogenous reactions: chemical reactions in which the reactants and products have the same phase (solid, liquid, gas)

\[ \text{H}_2(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow 2\text{HCl}(\text{g}) \]

\[ 2\text{KOH}(\text{l}) + \text{H}_2\text{C}_2\text{O}_4(\text{l}) \rightarrow \text{K}_2\text{C}_2\text{O}_4(\text{l}) + \text{H}_2\text{O}(\text{l}) \]
b) Heterogeneous reactions: chemical reactions in which the reactants and products have more than one phase.

\[ \text{FeS}_\text{(s)} + \text{HCl}_\text{(l)} \rightarrow \text{FeCl}_2\text{(l)} + \text{H}_2\text{S}_\text{(g)} \]

**Mass action law**: The rate of chemical reaction is directly proportional with formula conc. of reaction substances each raise to the power indicated by the number of ion or molecule appearing in the balance equation of the reaction.

**Equilibrium constant (K)**: A numerical quantity that relate the concentration of reactants and products in a chemical reaction to one another.

For chemical reaction:

\[ mA + nB \rightleftharpoons pC + bD \]

\[ V_f \propto [A] + [B] \]

\[ V_f = K [A] + [B] \]

\[ V_b \propto [C] + [D] \]

\[ V_b = K [C] + [D] \]

At equilibrium state:

\[ K [A][B] = K [C][D] \]

\[ \frac{K}{K} = K \frac{[C][D]}{[A][B]} \]

\[ K = \text{equilibrium constant} \]

\[ K = \frac{[C]^p[D]^b}{[A]^m[B]^n} \]
**Le chatelier principle**: The position of chemical equilibrium will always shift in a direction that tends to relieve the effect of an applied stress.

**Factors effecting the reactions at equilibrium**: 

1) *nature of reactant substances*: depends upon the difference in crystal and molecular structure.
   
   *Na* is react much faster than *Mg* with water.
   
   *Red phosphorus* is react much faster than *carbon*.

2) *Temperature*: the speed of chemical reactions will increase with the increasing of temperature. There are two kinds of thermal reactions:

   a) Exothermic reaction \( \Delta H(-) \)
   
   b) Endothermic reaction \( \Delta H(+) \)

   # Increasing of temperature favour forward reaction.
   # Decreasing of temperature favour backward reaction.

3) *Concentration*: The rate of chemical reaction will increase with the increasing of concentration.

4) *Solution and surface*: Increasing the surface of the solution due to an increasing in rate of reaction.

5) *Catalyste*: There are two kinds of catalysts:

   a) Positive catalyst; increase the rate of reaction.
   
   b) Negative catalyst; decrease the rate of reaction.

6) *Pressure*: it use when we deals with gas:

\[
\text{CO} + 2\text{H}_2 \rightleftharpoons 2\text{CH}_3\text{OH}
\]
Increasing of pressure will shift the reaction direction of less or lower volume this means, to forward direction.

While; \[ \text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightleftharpoons 2\text{HI} \]

Increasing in pressure will not effect on this reaction because the volume of product equal to volume of reactant.

**Equilibrium constant expressions:**

\( K_w = \) ion product (or ionization) constant for water.
\( K_{S.P} = \) solubility product constant.
\( K_a = \) ionization (or dissociation) constant of a weak acid.
\( K_b = \) ionization (or dissociation) constant of a weak base.
\( K_h = \) hydrolysis constant.
\( K_f = \) formation constant (mostly use for complex).
\( K_{\text{inst.}} = \) instability constant of complex ion.

**Ion product constant for water (Kw):**

Aqueous solutions contains small amount of hydronium \([\text{H}_3\text{O}^+]\) and hydroxide \([\text{OH}^-]\) ions as a consequence of the dissociation reaction:

\[
2\text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{OH}^- \quad \text{or} \quad \text{H}_2\text{O} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{OH}^- \\
\text{(acid)} \quad \text{(base)}
\]

\( \text{H}_2\text{O} \) acid gives \( \text{H}^+ \) to \( \text{H}_2\text{O} \) base to produce \( \text{H}_3\text{O}^+ \).

\[
K = \frac{[\text{H}_3\text{O}^+][\text{OH}^-]}{[\text{H}_2\text{O}]^2}
\]

\( K \ [\text{H}_2\text{O}]^2 = [\text{H}_3\text{O}^+][\text{OH}^-] = K_w \)

\( K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1 \times 10^{-14} \text{ mol}^2/\text{L}^2 \)
Kw is increase with the increasing of temperature, and it decrease with decreasing of temperature.

Variation of Kw with temperature:

<table>
<thead>
<tr>
<th>Temp. C°</th>
<th>Kw</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.0</td>
<td>0.114 x 10^{-14}</td>
</tr>
<tr>
<td>25</td>
<td>1.01 x 10^{-14}</td>
</tr>
<tr>
<td>50</td>
<td>5.47 x 10^{-14}</td>
</tr>
<tr>
<td>100</td>
<td>49 x 10^{-14}</td>
</tr>
</tbody>
</table>

Kw is used only for water.

Kw is applicable only for reversible reactions.

**Ex.:** Calculate hydronium and hydroxide ion conc. of pure water at 25C° and 100C°?

Because OH⁻ and H₃O⁺ are formed from the dissociation of water only, then their conc. must be equal, then:

\[ [H_3O^+] = [OH^-] \]

Substitution into equation (2-10) gives:

\[ [H_3O^+]^2 = [OH^-]^2 = Kw \]

\[ [H_3O^+] = [OH^-] = \sqrt{Kw} \]

At 25C° \[ [H_3O^+] = [OH^-] = \sqrt{1.01 \times 10^{-14}} = 1.01 \times 10^{-7} \]

At 100 C° \[ [H_3O^+] = [OH^-] = \sqrt{49 \times 10^{-14}} = 7.0 \times 10^{-7} \]
**EX.** Calculate hydronium and hydroxide ion conc. in 0.2M aqueous NaOH?

\[ \text{NaOH} \rightarrow \text{Na}^+ + \text{OH}^- \]

\[
\begin{array}{c}
0.2 \\
0.2 \\
0.2 \\
\end{array}
\]

\[ \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{OH}^- \]

\[ [\text{H}_3\text{O}^+] = [\text{OH}^-] = \text{Kw} \]

\[ [\text{H}_3\text{O}^+] = \frac{\text{Kw}}{[\text{OH}^-]} = \frac{1 \times 10^{-14}}{0.2} = 5 \times 10^{-14} \]

\[ [\text{OH}^-] = 0.2 + 5 \times 10^{-14} \]

**Equilibrium involving slightly soluble ionic solids:**

Most sparingly soluble salts are essentially completely dissociated in saturated aqueous solution.

\[ \text{AgCl}_\text{(s)} \rightleftharpoons \text{AgCl}_\text{(aq)} \]

\[ \text{AgCl}_\text{(aq)} \rightleftharpoons \text{Ag}^+ + \text{Cl}^- \]

\[ K_1 = \frac{[\text{AgCl}_\text{(aq)}]}{[\text{AgCl}_\text{(s)}]} \]

\[ K [\text{AgCl}_\text{(s)}] = [\text{AgCl}_\text{(aq)}] = K_1 \]

\[ K_2 = \frac{[\text{Ag}][\text{Cl}]}{[\text{AgCl}_\text{(aq)}]} \]
\[ K_1K_2 = [\text{AgCl}_{(aq)}] \times \frac{[\text{Ag}][\text{Cl}]}{[\text{AgCl}_{(aq)}]} \]

K.S.P = [Ag⁺][Cl⁻]

Where is:

K.S.P = solubility product constant that applied only for saturated solution.

**EX.** : How many grams of Ba(IO₃)₂ can be dissolved in 500 ml of water at 25°C? K.S.P Ba(IO₃)₂ = 1.57 x 10⁻⁹, M.wt = 487 gm / mole.

\[ \text{Ba(IO}_3\text{)}_2 \rightleftharpoons \text{Ba}^{2+} + 2\text{IO}_3^- \]

\[ x \quad 2x \]

K.S.P = [Ba⁺][IO₃⁻]²

K.S.P = (x)(2x)²

1.57 x 10⁻⁹ = 4x³

\[ X = 7.32 \times 10^{-4} \text{ mole/L or (M)} \]

Grams of \( \text{Ba(IO}_3\text{)}_2 \) = \( 7.32 \times 10^{-4} \text{ mmole/ml \times 500 ml \times 0.487 \text{ g/m mole} } \)

= 0.178 gm / 500 ml

**Dissociation equilibria for weak acids and basis:**

When a weak acids and basis dissolved in water, partial dissociation occurs.

For nitrus acid:

\[ \text{HNO}_2 + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{NO}_2^- \]

\[ \text{Ka} = \frac{[\text{H}_3\text{O}^+][\text{NO}_2^-]}{[\text{HNO}_2]} \]
Where $K_a$ is the acid dissociation constant.

To calculate $[\text{H}_3\text{O}^+]$ concentration for weak acid:

\[
\text{HA} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{A}^-
\]

\[
K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}
\]

$[\text{H}_3\text{O}^+] = [\text{A}^-]$

\[
K_a = \frac{[\text{H}_3\text{O}^+]^2}{[\text{HA}]}
\]

When dissociation is very weak:

Then, $[\text{HA}] = C_a$

Since, $[\text{H}_3\text{O}^+] = [\text{HA}^-]$

Therefore, $K_a = \frac{[\text{H}_3\text{O}^+]^2}{[\text{Ca}]}$

$[\text{H}_3\text{O}^+] = \sqrt{K_aC_a}$

For base:

\[
\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-
\]

\[
K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}
\]

When dissociation is very weak:

$[\text{NH}_3] = C_b$

Since, $[\text{OH}^-] = [\text{NH}_4]$

Therefore, $K_b = \frac{[\text{OH}^-]^2}{C_b}$
EX. : Calculate \([\text{H}_3\text{O}^+]\) concentration of an aqueous 0.120M nitrous acid solution? \(K_a = 5.1 \times 10^{-4}\).

\[
\text{HNO}_2 + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{NO}_2^-
\]

\[
K_a = \frac{[\text{H}_3\text{O}^+][\text{NO}_2^-]}{[\text{HNO}_2]} = 5.1 \times 10^{-4}
\]

\[
[\text{H}_3\text{O}^+] = [\text{NO}_2^-]
\]

\[
[\text{NO}_2^-] = 0.120 - [\text{H}_3\text{O}^+]
\]

\[
\frac{[\text{H}_3\text{O}^+]^2}{0.120} = 5.1 \times 10^{-4}
\]

\[
[\text{H}_3\text{O}^+] = \sqrt{0.120 \times 10^{-4}}
\]

EX. : calculate \([\text{H}_3\text{O}^+]\) concentration of an aqueous 0.2M of \(\text{NH}_3\)?

\(K_b = 1.8 \times 10^{-5}\) ?

\[
\text{NH}_3^+ + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-
\]

\[
K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3^+]}\]

\[
[\text{OH}^-] = [\text{NH}_4^+]
\]

\[
[\text{NH}_3^+] = 0.2 - [\text{OH}^-]
\]

\[
[\text{NH}_3] = \text{Cb}
\]

\[
1.8 \times 10^{-5} = \frac{[\text{OH}^-]^2}{\text{Cb}}
\]
\[ [\text{OH}^-] = \sqrt{K_{b}C_{b}} = \sqrt{1.8 \times 10^{-5} \times 0.2} \]
\[ [\text{H}_3\text{O}^+] = \frac{K_{w}}{[\text{OH}^-]} \]
\[ [\text{H}_3\text{O}^+] = \frac{1 \times 10^{-14}}{\sqrt{1.8 \times 10^{-5} \times 0.2}} \]

Relationship between dissociation constant for conjugate acid-base pair:

\[ \text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^- \]

\[ K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} \quad \ldots \quad (1) \]

\[ \text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{NH}_3 + \text{H}_3\text{O}^+ \]

\[ K_a = \frac{[\text{H}_3\text{O}^+][\text{NH}_3]}{[\text{NH}_4^+]} \quad \ldots \quad (2) \]

Multiplication of equation (1) and (2) gives:

\[ K_a K_b = \frac{[\text{H}_3\text{O}^+][\text{NH}_3]}{[\text{NH}_4^+]} \times \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} \]

\[ K_a K_b = [\text{H}_3\text{O}^+][\text{OH}^-] \]

but, \[ [\text{H}_3\text{O}^+][\text{OH}^-] = K_w \]

Therefore,

\[ K_a K_b = K_w \]

**EX.** Calculate \( K_b \) value for the reaction \( \text{CN}^- + \text{H}_2\text{O} \rightleftharpoons \text{HCN} + \text{OH}^- \) when \( K_{\text{HCN}} = 2.1 \times 10^{-9} \)?

\[ K_b = \frac{[\text{HCN}][\text{OH}^-]}{[\text{CN}^-]} = \frac{K_w}{K_{\text{HCN}}} \]

\[ K_b = \frac{1 \times 10^{-14}}{2.1 \times 10^{-9}} = 4.8 \times 10^{-6} \]
**Complex formation equilibria** \( K_f \):

One of important class of reactions involves the formation of soluble complex ions :

\[
Fe^{3+} + SCN^- \rightleftharpoons Fe(\text{SCN})^{2+}
\]

\[
K_f = \frac{[Fe(\text{SCN})^{2+}]}{[\text{SCN}^-][Fe^{3+}]}
\]

\[
K_{\text{inst.}} = \frac{[\text{SCN}^-][Fe^{3+}]}{[Fe(\text{SCN})]}
\]

\[
K_f \times K_{\text{inst.}} = 1 \quad \text{or} \quad K_f = \frac{1}{K_{\text{inst.}}}
\]

**Stepwise equilibrium** :

Many weak electrolites associate or dissociate in a stepwise manner, and an equilibrium constant can be written for each step, when ammonia is added to a solution containing silver ions :

\[
\text{Ag}^+ + \text{NH}_3 \rightleftharpoons \text{AgNH}_3^+
\]

\[
K_1 = \frac{[\text{AgNH}_3^+]}{[\text{NH}_3][\text{Ag}]}
\]

\[
\text{AgNH}_3^+ + \text{NH}_3 \rightleftharpoons \text{Ag(NH}_3)_2
\]

\[
K_2 = \frac{[\text{Ag(NH}_3)_2]}{[\text{NH}_3][\text{Ag(NH}_3)]}
\]

\[
K_f = K_1 \times K_2 = \frac{[\text{Ag(NH}_3)_2]}{[\text{NH}_3][\text{Ag}]}
\]

For dissociation :

\[
H_3\text{PO}_4 + H_2\text{O} \rightleftharpoons H_2\text{PO}_4^- + H_3\text{O}^+
\]

\[
K_{a_1} = \frac{[H_2\text{PO}_4^-][H_3\text{O}^+]}{[H_3\text{PO}_4]}
\]
\[ \text{H}_2\text{PO}_4^{-2} + \text{H}_3\text{O}^+ \rightleftharpoons \text{HPO}_4^{-2} + \text{H}_3\text{O}^+ \]

\[ \text{Ka}_2 = \frac{[\text{H}_2\text{PO}_4^{-2}][\text{H}_3\text{O}^+]}{[\text{H}_2\text{PO}_4^{-2}]} \]

\[ \text{HPO}_4^{-2} + \text{H}_3\text{O}^+ \rightleftharpoons \text{PO}_4^{-3} + \text{H}_3\text{O}^+ \]

\[ \text{Ka}_3 = \frac{[\text{PO}_4^{-3}][\text{H}_3\text{O}^+]}{[\text{H}_2\text{PO}_4^{-2}]} \]
**CHAPTER TWO**

**Solutions**: Homogenous mixture for two or more substances, one of them called solute and the other one is the solvent.

- The solute is an active substance in reaction and the solvent is active.
- There are many solutions formed by dissolving gases in liquids like CO\textsubscript{2} in water.
- Homogenous solution: Partial homogenous mixing for two or more substances that will not react chemically with each other.
- Heterogenous solution: Represent with suspended and colloidal solution.
- Aqueous solution: is the solution in which water is the solvent.
- Nonaqueous solution: when the solvent is any substance except water.

Classification of solutions are depends upon:

a) *nature of molecules* (or the volume of atoms or molecules of solute), these solutions are:
   1) True solution.
   2) Suspended solution.
   3) Colloidal solution.

b) *solute concentration in solution*, it can be classified into:
   1) Saturated solution.
   2) Unsaturated solution.
3) super saturated solution.

**Units of weight and concentration:**

Concentration of solution: is the weight of solute dissolved in a given amount (volume or weight) of solvent.

Weight of solute can be expressed by:

1) physical units, e.g. gram (g), milligram (mg), kilogram (kg).

2) chemical units: includes,
   a) g- atomic weight.
   b) g- molecular weight.
   c) g- formula weight.
   d) g- ionic weight.
   e) g- equivalent weight.

**a) g- atomic weight (g- A.wt):**

A chemical unit in which use only for elements (or atom):

\[
g\text{-atom} = \frac{\text{wt(gm)}}{\text{A.wt}(\frac{g}{\text{g-atom}})}
\]

**EX.:** calculate the no. of grams of \( \text{O}_2 \) in 5 g-atom?

\[
g\text{-atom} = \frac{\text{wt(gm)}}{\text{A.wt}(\frac{g}{\text{g-atom}})}
\]

\[
\text{wt(gm)} = \text{g-atom} \times \text{A.wt}
\]

\[
\text{wt(gm)} = 5 \text{ g-atom} \times 32 \frac{g}{\text{g-atom}} = 160 \text{ gm}
\]

**b) g-molecular weight (g-M.wt):**
A chemical unit of weight used for substance found as molecules like, NH₃, H₂O, ........... ect.

\[
g\text{-M.wt} = \frac{\text{wt(gm)}}{M\text{.wt}\left(\frac{g}{\text{mol}}\right)}
\]

**EX.** Calculate no. of grams of H₂O that is found in 3 mole of molecule?

\[
g\text{- M.wt} = \frac{\text{wt(gm)}}{M\text{.wt}\left(\frac{g}{\text{mol}}\right)}
\]

\[
\text{wt. H₂O} = \text{mole} \times M\text{.wt}\left(\frac{g}{\text{mol}}\right)
\]

\[
\text{wt. H₂O} = 3 \text{ mole} \times 18\left(\frac{g}{\text{mol}}\right) = 54 \text{ gm}
\]

**C) g-formula wieght(g- fw.) :**

A chemical unit of weight used for chemical compounds that consider as strong electrolite like; HCl, HNO₃, KOH ........... ect.

# molecular weight have the same value for all the compounds.

**d) g-ionic weight (g-ion) :**

A chemical unit of weight used for ions (anions and cations).

**EX.** Aqueous solution contain 14.2 gm of Na₂SO₄, calculate how many g-ion of Na⁺ and SO₄⁻² in solution, M.wt of Na₂SO₄ = 142 g/mole?

\[
\text{Na₂SO₄} \rightarrow 2\text{Na}^+ + \text{SO}_4^{-2}
\]

\[
\text{No. of moles} = \frac{\text{wt(gm)}}{M\text{.wt}\left(\frac{g}{\text{mol}}\right)}
\]
No. of moles = \( \frac{14.2 \text{ g}}{142 \left( \frac{\text{g}}{\text{mol}} \right)} \) = 0.1 mole

g-ion of Na\(^+\) = 2 \times 0.1 = 0.2 g-ion.

g-ion of SO\(_4\)^{2-} = 1 \times 0.1 = 0.1 g-ion.

**EX.** Calculate the no. of gram of Cl\(^-\) in 4 g-ion of Cl\(^-\)?

\[
= 142 \text{ gm} \quad 4 \text{ g-ion} \times 35.5 \frac{\text{gm}}{\text{g-ion}}
\]

**Mole:** is the summation of the atomic weight in grams for all of the atoms in the chemical formula for the species, or it equal to Avogadro’s number (6 \times 10^{23}) of atoms, ions, and molecules.

No. of moles = \( \frac{\text{wt(gm)}}{\text{M.wt} \left( \frac{\text{g}}{\text{mol}} \right)} \)

**Methods for the expression of concentration of solutions:**

There are several methods for express the concentration of solutions:

1) **Molarity (M):** moles of solute contained in one liter of solution or millimole in one milliliter.

\[
M = \frac{\text{mole}}{\text{liter}} \quad \text{or} \quad \frac{\text{m.mole}}{\text{ml}}
\]

\[
M = \frac{\text{no. of mole of solute}}{\text{liter of solution}} = \frac{\text{wt(gm)}}{\text{M.wt} \left( \frac{\text{g}}{\text{mol}} \right)} = \frac{\text{mole}}{\text{liter}}
\]

**EX.** Calculate the molarity of 0.2 gm of NaOH in one liter of solution, M.wt of NaOH = 40 gm / mole?
\[
M = \frac{\text{mole}}{\text{liter}} = \frac{\text{wt(gm)}}{\text{M.wt(gm/mole)}}
\]

\[
= \frac{0.2 \text{(gm)}}{40 \text{(gm/mole)}} \times \frac{1 \text{ L}}{1 \text{ L}} = 0.005 \frac{\text{mole}}{\text{liter}}
\]

**EX.** Prepare 0.1 M of NaOH in one liter of solution, M.wt. of NaOH = 40 gm / mole?

\[
M = \frac{\text{wt(gm)}}{\text{M.wt(gm/mole)}}
\]

\[
0.1 = \frac{\text{wt(gm)}}{40 \text{(gm/mole)}} \times \frac{1 \text{ liter}}{1 \text{ liter}}
\]

Wt. = 0.1 x 40 x 1 = 0.4 gm

**In liquid substance:**

\[
M = \frac{\text{sp.gra.} \times \% \times 1000}{\text{M.wt}}
\]

Where is: sp.gra. = specific gravity.

\% = gm or vol. of solute in 100ml of solvent.

Also: \[
M = \frac{\text{wt}}{\text{M.wt}} \times \frac{1000}{V \text{(ml)}}
\]

**EX.** Calculate the M of NaOH solution that prepared by dissolving 0.1 gm in 500 ml of water, M.wt of NaOH = 40 gm / mole?

\[
M = \frac{\text{wt}}{\text{M.wt}} \times \frac{1000}{V \text{(ml)}}
\]
$$M = \frac{0.1 \text{ gm}}{40 \text{ gm/mol}} \times \frac{1000}{500 \text{ ml}}$$

$$M = 0.005 \text{ mol / liter}$$

# In case of dilution we can use dilution low:

$$M_1 V_1 = M_2 V_2$$

$$\frac{\text{mole}_1}{\text{liter}_1} \cdot \text{liter}_1 = \frac{\text{mole}_2}{\text{liter}_2} \cdot \text{liter}_2$$

Therefore, $$\text{mole}_1 = \text{mole}_2$$

**EX.** Prepare 250 ml of 0.1M from 2M concentrated H$_2$SO$_4$, M.wt = 98 gm/mol?

$$M_1 V_1 = M_2 V_2$$

$$2 \times V_1 = 0.1 \times 250$$

$$V_1 = 12.5 \text{ ml}$$

That means, we should take 12.5 ml from concentrated H$_2$SO$_4$ and dissolve it in water and complete the volume to 250 ml in volumetric flask.

2) **Normality (N):** number of g-equivalent weight of solute that are contained in a liter of solution.

$$N = \frac{\text{no. of g–eq.wt of solute}}{\text{liter of solution}}$$

$$\text{g- eq.wt.} = \frac{\text{wt(gm)}}{\text{eq.wt}}$$
\[ \text{eq.wt.} = \frac{\text{M.wt}}{n} \]

\[ N = \frac{\frac{\text{wt(gm)}}{\text{eq.wt}}\left(\text{gm}\right)}{\text{liter}} = \frac{\text{eq.wt}}{\text{liter}} \]

**Note**: 
1) no. of H⁺. 
2) no. of OH⁻. 
3) no. of charges (negative or positive). 
4) no. of electrons.

**EX.** Calculate \( N \) of 0.2 gm of \( \text{H}_2\text{SO}_4 \) in one liter of solution, M.wt of \( \text{H}_2\text{SO}_4=98 \text{gm/mole} \) ?

\[ \frac{\text{wt(gm)}}{\text{eq.wt}}\left(\text{gm}\right) \]

\[ N = \frac{\frac{98 \text{gm}}{2}}{\text{liter}} = 49 \]

\[ \frac{\frac{0.2(\text{gm})}{49(\text{eq.wt})}}{1 \text{ liter}} = 0.004 \text{ eq.wt.} / \text{liter} \]

Also, \( N = \frac{\text{wt}}{\text{eq.wt}} \times \frac{1000}{V \text{ (mL)}} \)

\[ N = \frac{\text{sp.gra.} \times \% \times 1000}{\text{eq.wt}} \]

\( N_1 \ V_1 = N_2 \ V_2 \) (in case of dilution).
General principles to calculate the equivalent weight:

a) for acids:

\[ \text{eq.wt of acid} = \frac{\text{M.wt of acid}}{\text{no.of H+ active}} \]

There are many types of acids:

1) mono basic acid: \( \text{HCl, HNO}_3, \text{HF} \), ........... ect.
2) di basic acid: \( \text{H}_2\text{SO}_4, \text{H}_2\text{S}, \text{H}_2\text{SO}_3 \), ...........ect.
3) tri basic acid: \( \text{H}_3\text{PO}_4 \), .........................ect.

b) for base:

\[ \text{eq.wt of acid} = \frac{\text{M.wt of base}}{\text{no.of OH− active}} \]

\[ \text{HCl} + \text{NH}_4\text{OH} \rightarrow \text{NH}_4\text{Cl} + \text{H}_2\text{O} \]

\[ \text{eq.wt. of NH}_4\text{OH} = \frac{\text{M.wt of NH}_4\text{OH}}{\text{no.of OH− active}} = \frac{35}{1} = 35 \]

\[ \text{Ca(OH)}_2 + \text{H}_2\text{SO}_4 \rightarrow \text{CaSO}_4 + \text{H}_2\text{O} \]

\[ \text{eq.wt.of Ca(OH)}_2 = \frac{\text{M.wt of Ca(OH)}_2}{\text{no.of OH− active}} = \frac{74}{1} = 74 \]

c) Eq.wt. for substances reacts with acid or base:
SO₃ + H₂O → H₂SO₄

\[
eq wt. \text{ of } SO_3 = \frac{M.wt \text{ of } SO_3}{2}
\]

Na₂O + 2HCl → 2NaCl + H₂O

\[
eq wt. \text{ of } Na_2O = \frac{M.wt \text{ of } Na_2O}{2}
\]

d) for salts:

\[
eq wt. \text{ of salt} = \frac{M.wt \text{ of salt}}{\text{no. of positive ion or no. of negative ion}}
\]

Ag⁺NO₃⁻ + H⁺Cl⁻ → Ag⁺Cl⁻ + H⁺NO₃⁻

\[
eq wt. \text{ of } Ag^+NO_3^- = \frac{M.wt \text{ of } Ag+NO_3^-}{1}
\]

\[
eq wt. \text{ of } Ag_2^+O^{2-} = \frac{M.wt}{2}
\]

\[
eq wt. \text{ of } Al^{+3}PO_4^{-3} = \frac{M.wt}{3}
\]

\[
eq wt. \text{ of } Ca_3^{+2}(PO_4)_2^{-3} = \frac{M.wt}{6}
\]

e) for oxidation-reduction reaction:

In these reactions, the equivalent weight is equal to atomic, ionic, or molecular weight divided into the no. of losted or accepted electrons in one molecule.

\[
eq wt. \text{ of reducing agent} = \frac{M.wt}{\text{no. of electrons losted for one molecule}}
\]
eq.wt. of oxidizing agent = \frac{M.wt}{\text{no. of electrons accepted for one molecule}}

\[ \text{KMn}^{7+}\text{O}_4 + \text{Fe}^{2+} \text{SO}_4^{2-} + \text{H}_2\text{SO}_4 \rightarrow \text{Mn}^{2+}\text{SO}_4^{2-} + \text{H}_2\text{O} + \text{Fe}^{2+3}\text{(SO}_4)_3 \]

\[ \text{Mn}^{7+}\text{O}_4^- + \text{H}^+ + 5e^- \rightarrow \text{Mn}^{2+} + \text{H}_2\text{O} \quad \text{.............. In acidic medium} \]

eq.wt. Mn\(^{7+}\text{O}_4^- = \frac{M.wt}{5} \]

\[ \text{Mn}^{7+}\text{O}_4^- + \text{H}_2\text{O} \rightarrow \text{Mn}^{4+}\text{O}_2^- + \text{OH}^- \quad \text{.............. In basic medium} \]

eq.wt. Mn\(^{7+}\text{O}_4^- = \frac{M.wt}{3} \]

f) for complex ion:

eq.wt. of complex ion = \frac{M.wt}{\text{no. of charges}}

\[ \text{Ag} + 2\text{NH}_3 \rightarrow \text{Ag}^{+}(\text{NH}_3)_2^- \]

eq.wt. of \( \text{Ag}^{+}(\text{NH}_3)_2^- = \frac{M.wt}{1} \]

\[ \text{Cu}^{+2} + 4\text{NH}_3 \rightarrow \text{Cu}(\text{NH}_3)_4^{-2} \]

eq.wt. of \( \text{Cu}(\text{NH}_3)_4^{-2} = \frac{M.wt}{2} \]

3) part per million (ppm): used for very dilute solution:

\[ \text{ppm} = \frac{\text{wt of solute (mg)}}{\text{vol. of solution (L)}} \quad \text{or} \quad \frac{\text{mg of solute}}{10^6} \quad \text{or} \quad \frac{\text{mg of solute}}{\text{liter of solution}} \]

\[ \text{ppm} = \frac{\text{wt of solute (mg)}}{\text{wt.of solution (L)}} \times 10^6 \]
ppm = mg / L  or $\mu$g / ml

**EX.** 1 gm of solute dissolved in 1 liter solution, calculate Cppm?

1 gm of solute must convert to mg.

1 gm x 1000 = 1000 mg

1000 mg / 1L solution = 1 ppm

Or 1 gm convert to $\mu$g

$\therefore$ 1x1000000 = 1000000 $\mu$g

Also, 1 liter solution convert to ml

1L solution x1000 = 1000 ml

$\therefore$ ppm = 1000000 $\mu$g / 1000 ml = 1000 ppm

**4) percentage concentration:**

There are several methods for express concentration in terms of percentage (part per hundred), these methods includes:

a) weight percentage (w/w) = \[
\frac{\text{wt of solute}}{\text{wt. of solution}} \times 100
\]

weight percentage is used to express the concentration of commercial aqueous solution or reagents. For example, on nitric acid bottle we find 70% solution, that means the reagent contains 70 gm of nitric acid per 100gm of solution.

b) volume percentage (v/v) = \[
\frac{\text{vol of solute}}{\text{vol. of solution}} \times 100
\]

volume percentage is used when we need to dilute concentrated liquid by another solvent to prepare a dilute solution. For example, 10% HCl solution, that mean you should take 10ml of concentrated HCl and dissolving in sufficient water to give 100ml of solution.
c) weight-volume percentage = \( \frac{\text{wt. of solute (g)}}{\text{vol. of solution (ml)}} \times 100 \)

weight-volume percentage is used when we need to prepare a dilute aqueous solution of solid reagents. For example, 5% aqueous silver nitrate is prepared by dissolving 5gm of silver nitrate in sufficient water to give 100ml of solution.