FORMULA BOOK SEM 1

Methods of expressing concentration of solution

Concentration of solution is the amount of solute dissolved in a known amount of the solvent or solution. The concentration of solution can be expressed in various ways as discussed below,

(1) Percentage : It refers to the amount of the solute per 100 parts of the solution. It can also be called as parts per hundred (*pph*). It can be expressed by any of following four methods,

(i) *Weight to weight percent*

% *w/w* $= \dfrac{\text{Wt. of solute}}{\text{Wt. of solution}} \times 100$

Example : **10%** Na_2CO_3 **solution** *w/w* means $10g$ of Na_2CO_3 is dissolved in $100g$ of the solution. (It means $10g$ Na_2CO_3 is dissolved in $90g$ of H_2O)

(ii) *Weight to volume percent*

% *w/v* $= \dfrac{\text{Wt. of solute}}{\text{Volume of solution}} \times 100$

Example : **10%** Na_2CO_3 (*w/v*) means $10g$ Na_2CO_3 is dissolved in $100cc$ of solution.

(iii) *Volume to volume percent*

% *v/v* $= \dfrac{\text{Vol. of solute}}{\text{Vol. of solution}} \times 100$

Example : **10% ethanol** (*v/v*) means $10cc$ of ethanol dissolved in $100cc$ of solution.

(iv) *Volume to weight percent*

% *v/w* $= \dfrac{\text{Vol. of solute}}{\text{Wt. of solution}} \times 100$

Example : **10% ethanol** (*v/w*) means $10cc$ of ethanol dissolved in $100g$ of solution.

(2) Parts per million (*ppm*) and parts per billion (*ppb*) : When a *solute is present in trace quantities*, it is convenient to express the concentration in parts per million and parts per billion. It is the

number of parts of solute per million (10^6) or per billion (10^9) parts of the solution. *It is independent of the temperature*.

$ppm = \dfrac{\text{mass of solute component}}{\text{Total mass of solution}} \times 10^6$

$$ppb = \frac{\text{mass of solute component}}{\text{Total mass of solution}} \times 10^9$$

(3) Strength : The strength of solution is defined as the amount of solute in grams present in one *litre* (or dm^3) of the solution. It is expressed in *g*/*litre* or (g/dm^3).

$$\text{Strength} = \frac{\text{Mass of solute in grams}}{\text{Volume of solution in litres}}$$

(4) Normality (N) : It is defined as the number of gram equivalents (equivalent weight in grams) of a solute present per *litre* of the solution. Unit of normality is gram equivalents *litre*-1. Normality changes with temperature since it involves volume. When a solution is diluted x times, its normality also decreases by x times. Solutions in term of normality generally expressed as,

N = **Normal solution;** $5N$ = **Penta normal,**

$10N$ = **Deca normal;** $N/2$ = **semi normal**

$N/10$ = **Deci normal;** $N/5$ = **Penti normal**

$N/100$ **or** $0.01\ N$ = **centinormal,**

$N/1000$ **or** **0.001= millinormal**

Mathematically normality can be calculated by following formulas,

(i) $\text{Normality}(N) = \dfrac{\text{Number of g.eq. of solute}}{\text{Volume of solution}(l)}$

(ii) $N = \dfrac{\text{Weight of solute in } g.}{g.\text{eq. weight of solute} \times \text{Volume of solution}(l)}$

(iii) $N = \dfrac{\text{Wt. of solute per litre of solution}}{g\ \text{eq. wt. of solute}}$

(iv) $N = \dfrac{\text{Wt. of solute}}{g.\text{eq. wt. of solute}} \times \dfrac{1000}{\text{Vol. of solution in } ml}$

(v) $N = \dfrac{\text{Percent of solute} \times 10}{g\ \text{eq. wt. of solute}}$,

(vi) $N = \dfrac{\text{Strength in } g\ l^{-1} \text{ of solution}}{g\ \text{eq. wt. of solute}}$

(vii) $N = \dfrac{\text{Wt\% } \times \text{density} \times 10}{\text{Eq. wt.}}$

(viii) If volume V_1 and normality N_1 is so changed that new normality and volume N_2 and V_2 then,

$$N_1 V_1 = N_2 V_2 \text{ (Normality equation)}$$

(ix) When two solutions of the same solute are mixed then normality of mixture (N) is

$$N = \frac{N_1 V_1 + N_2 V_2}{V_1 + V_2}$$

(x) Vol. of water to be added i.e., $(V_2 - V_1)$ to get a solution of normality N_2 from $V_1\ ml$ of normality N_1

$$V_2 - V_1 = \left(\frac{N_1 - N_2}{N_2} \right) V_1$$

(xi) If $W\ g$ of an acid is completely neutralised by $V\ ml$ of base of normality N

$$\frac{\text{Wt. of acid}}{g\ \text{eq. wt. of acid}} = \frac{VN}{1000}$$

Similarly, $\dfrac{\text{Wt. of base}}{g\ \text{eq. wt. of base}} = \dfrac{\text{Vol. of acid} \times N \text{ of acid}}{1000}$

(xii) When $V_a\ ml$ of acid of normality N_a is mixed with $V_b\ ml$ of base of normality N_b

(a) If $V_a N_a = V_b N_b$ **(Solution neutral)**

(b) If $V_a N_a > V_b N_b$ **(Solution is acidic)**

(c) If $V_b N_b > V_a N_a$ **(Solution is basic)**

(xiii) Normality of the acidic mixture $= \dfrac{V_a N_a + V_b N_b}{(V_a + V_b)}$ **(xiv) Normality of the basic mixture** $= \dfrac{V_b N_b + V_a N_a}{(V_a + V_b)}$

(xv) $N = \dfrac{\text{No. of meq * of solute}}{\text{Vol. of solution in } ml}$ **(* 1 equivalent = 1000 milliequivalent or meq.)**

(5) Molarity (M) : Molarity of a solution is the *number of moles of the solute per litre of solution (or number of millimoles per ml. of solution)*. Unit of molarity is **mol/litre** or **mol/dm³** For example, a molar $(1M)$ solution of sugar means a solution containing 1 mole of sugar (*i.e.*, 342 *g* or 6.02×10^{23} molecules of it) per *litre* of the solution. Solutions in term of molarity generally expressed as,

$1M$ = Molar solution, $2M$ = Molarity is two,

$\dfrac{M}{2}$ or 0.5 *M* = Semimolar solution,

$\dfrac{M}{10}$ or 0.1 *M* = Decimolar solution,

$\dfrac{M}{100}$ or 0.01 *M* = Centimolar solution

$\dfrac{M}{1000}$ or 0.001 *M* = Millimolar solution

• Molarity is most common way of representing the concentration of solution.

• Molarity is depend on temperature as, $M \propto \dfrac{1}{T}$

• When a solution is diluted (*x* times), its molarity also decreases (by x times)
Mathematically molarity can be calculated by following formulas,

(i) $M = \dfrac{\text{No. of moles of solute} (n)}{\text{Vol. of solution in litres}}$,

(ii) $M = \dfrac{\text{Wt. of solute (in gm) per litre of solution}}{\text{Mol. wt. of solute}}$

(iii) $M = \dfrac{\text{Wt. of solute (in gm)}}{\text{Mol. wt. of solute}} \times \dfrac{1000}{\text{Vol. of solution in } ml.}$

(iv) $M = \dfrac{\text{No. of millimoles of solute}}{\text{Vol. of solution in } ml}$

(v) $M = \dfrac{\text{Percent of solute} \times 10}{\text{Mol. wt. of solute}}$

(vi) $M = \dfrac{\text{Strength in } gl^{-1} \text{ of solution}}{\text{Mol. wt. of solute}}$

(vii) $M = \dfrac{10 \times \text{Sp. gr. of the solution} \times \text{Wt.\% of the solute}}{\text{Mol. wt. of the solute}}$

(viii) If molarity and volume of solution are changed from M_1, V_1 to M_2, V_2. Then,

$$M_1 V_1 = M_2 V_2 \text{ (Molarity equation)}$$

(ix) In balanced chemical equation, if n_1 moles of reactant one react with n_2 moles of reactant two. Then,

$$\frac{M_1 V_1}{n_1} = \frac{M_2 V_2}{n_2}$$

(x)   If two solutions of the same solute are mixed then molarity (M) of resulting solution.

$$M = \frac{M_1 V_1 + M_2 V_2}{(V_1 + V_2)}$$

(xi) Volume of water added to get a solution of molarity M_2 from V_1 ml of molarity M_1 is

$$V_2 - V_1 = \left(\frac{M_1 - M_2}{M_2} \right) V_1$$

Relation between molarity and normality

Normality of solution = molarity $\times \dfrac{\text{Molecular mass}}{\text{Equivalent mass}}$

Normality $\times$ equivalent mass = molarity $\times$ molecular mass

For an acid, $\dfrac{\text{Molecular mass}}{\text{Equivalent mass}}$ = Basicity

So, Normality of acid = molarity $\times$ basicity.

For a base, $\dfrac{\text{Molecular mass}}{\text{Equivalent mass}}$ = Acidity

So, Normality of base = Molarity $\times$ Acidity.

(6) Molality (m) : *It is the number of moles or gram molecules of the solute per* **1000** *g of the solvent.* **Unit of molality is** mol/kg. **For example, a** 0.2 **molal** $(0.2m)$ **solution of glucose means a solution obtained by dissolving 0.2 mole of glucose in** $1000gm$ **of water.** *Molality (m) does not depend on temperature* **since it involves measurement of weight of liquids.** *Molal solutions are less concentrated than molar solution.*

Mathematically molality can be calculated by following formulas,

(i) $m = \dfrac{\text{Number of moles of the solute}}{\text{Weight of the solvent in grams}} \times 1000$

(ii) $m = \dfrac{\text{Strength per 1000 grams of solvent}}{\text{Molecular mass of solute}}$

(iii) $m = \dfrac{\text{No. of } gm \text{ moles of solute}}{\text{Wt. of solvent in } kg}$

(iv) $m = \dfrac{\text{Wt. of solute}}{\text{Mol. wt. of solute}} \times \dfrac{1000}{\text{Wt. of solvent in } g}$

(v) $m = \dfrac{\text{No. of millimoles of solute}}{\text{Wt. of solvent in } g}$

(vi) $m = \dfrac{10 \times \text{solubility}}{\text{Mol. wt. of solute}}$

(vii) $m = \dfrac{1000 \times \text{wt.\% of solute}(x)}{(100 - x) \times \text{mol. wt. of solute}}$

(viii) $m = \dfrac{1000 \times \text{Molarity}}{(1000 \times \text{sp. gravity}) - (\text{Molarity} \times \text{Mol.wt. of solute})}$

Relation between molarity (*M*) and molality (*m*)

Molality (*m*) = $\dfrac{\text{Molarity}}{\text{Density} - \dfrac{\text{Molarity} \times \text{molecular mass}}{1000}}$

Molarity (*M*) = $\dfrac{\text{Molality} \times \text{density}}{1 + \dfrac{\text{Molality} \times \text{molecular mass}}{1000}}$

(7) Formality (*F*) : Formality of a solution may be defined as the number of gram formula masses of the ionic solute dissolved per litre of the solution. It is represented by F. Commonly, the term formality is used to express the concentration of the ionic solids which do not exist as molecules but exist as network of ions. A solution containing one gram formula mass of solute per litre of the solution has formality equal to one and is called formal solution. It may be mentioned here that the *formality of a solution changes with change in temperature*.

Formality (F)= $\dfrac{\text{Number of gram formula masses of solute}}{\text{Volume of solution in litres}}$

= $\dfrac{\text{Mass of ionic solute}(g)}{(gm. \text{ formula mass of solute}) \times (\text{Volume of solution}(l))}$

Thus, $F = \dfrac{W_B(g)}{GFM \times V(l)}$ **or** $\dfrac{W_B(g) \times 1000}{GFM \times V(ml)}$

(8) Mole fraction (*X*) : *Mole fraction may be defined as the ratio of number of moles of one component to the total number of moles of all the components (solvent and solute) present in the solution*. It is denoted by the letter X. It may be noted that *the mole fraction is independent of the temperature. Mole fraction is dimensionless.* Let us suppose that a solution contains the components A and B and suppose that $W_A g$ of A and $W_B g$ of B are present in it.

Number of moles of A **is given by,** $n_A = \dfrac{W_A}{M_A}$ **and**

the number of moles of B **is given by,** $n_B = \dfrac{W_B}{M_B}$

where M_A and M_B are molecular masses of A and B respectively.

Total number of moles of A **and** $B = n_A + n_B$

Mole fraction of A **,** $X_A = \dfrac{n_A}{n_A + n_B}$

Mole fraction of B **,** $X_B = \dfrac{n_B}{n_A + n_B}$

The sum of mole fractions of all the components in the solution is always one.

$$X_A + X_B = \dfrac{n_A}{n_A + n_B} + \dfrac{n_B}{n_A + n_B} = 1.$$

Thus, if we know the mole fraction of one component of a binary solution, the mole fraction of the other can be calculated.

Relation between molality of solution (*m*) and mole fraction of the solute (X_A).

$$X_A = \dfrac{m}{55.5 + m}$$

(9) Mass fraction : Mass fraction of a component in a solution is the mass of that component divided by the total mass of the solution. For a solution containing $w_A gm$ of *A* and $w_B gm$ of *B*

Mass fraction of $A = \dfrac{w_A}{w_A + w_B}$ **; Mas fraction of** $B = \dfrac{w_B}{w_A + w_B}$

It may be noted that molality, mole fraction, mass fraction etc. are preferred to molarity, normality, etc. because the former involve the weights of the solute and solvent where as later involve volumes of solutions. Temperature has no effect on weights but it has significant effect on volumes.

(10) Demal unit (D) : The concentrations are also expressed in "Demal unit". One demal unit represents one mole of solute present in one litre of solution at $0^o C$.

(4) Avogadro's hypothesis : "Equal volumes of all gases under similar conditions of temperature and pressure contain equal number of molecules." Avogadro hypothesis has been found to explain as follows,

(i) Provides a method to determine the atomic weight of gaseous elements.

(ii) *Provides a relationship between vapour density (V.D.) and molecular masses of substances*.

Molecular mass $= 2 \times$ vapour density

(iii) *It helps in the determination of mass of fixed volume of a particular gas.*

Mass of $1\, ml$ **gas = V.D.** $\times$ **0.0000897** *gm.*

(iv) *It also helps in the determination of molar volume at N.T.P.*

$\because$ **V.D.** $\times$ **0.0000897** *gm.* **gas has volume =** $1\, ml$

$\therefore$ **2** $\times$ **V.D.(i.e., molecular mass)** *gm.* **has** $\text{volume} = \dfrac{1 \times 2 \times V.D.}{V.D. \times 0.0000897}\, ml = 22400\, ml$

$\therefore$ **Molar mass of a gas or its 1 mole occupies 22.4** *L* **volume at S.T.P.**

(v) It helps in determination of molecular formulae of gases and is very useful in gas analysis. By knowing the molecular volumes of reactants and products of reaction, molecular composition can be determined easily.

Atomic, Molecular and Equivalent masses

(1) Atomic mass : It is the average relative mass of atom of element as compared with an atom of carbon –12 isotope taken as 12.

$$\text{Atomic mass} = \frac{\text{Average mass of an atom}}{1/12 \times \text{Mass of an atom of } C^{12}}$$

Average atomic mass : If an elements exists in two isotopes having atomic masses '*a*' and '*b*' in the ratio $m : n$, then average atomic mass $= \dfrac{(m \times a) + (n \times b)}{m + n}$. Since the atomic mass is a ratio, it has no units and is expressed in *amu*, 1 *amu* $= 1.66 \times 10^{-24}\, g$. One atomic mass unit (amu) is equal to $\dfrac{1}{12} th$ of the mass of an atom of carbon-12 isotope.

Gram atomic mass (GAM) : Atomic mass of an element expressed in grams is called *Gram atomic mass or gram atom or mole atom*.

(i) Number of gram atoms = $\dfrac{\text{Mass of an element}}{\text{GAM}}$

(ii) Mass of an element in *gm.* **= No. of** *gm.* **atom** $\times$ **GAM**

(iii) Number of atoms in 1 GAM = 6.02 $\times 10^{23}$

$\therefore$ **Number of atoms in a given substance**

= No. of GAM $\times$ **6.02** $\times 10^{23}$ **=** $\dfrac{\text{Mass}}{\text{GAM}}$ $\times$ **6.02** $\times 10^{23}$

(iv) Number of atoms in 1 *gm* **of element** $= \dfrac{6.02 \times 10^{23}}{\text{Atomic mass}}$

(v) Mass of one atom of the element (in *gm.***)** $= \dfrac{\text{GAM}}{6.02 \times 10^{23}}$

Methods of determination of atomic mass

(i) Dulong and Pettit's method : According to Dulong and Pettit's law

Atomic mass $\times$ **Specific heat = 6.4 (***approx.***)**

Atomic mass (*approx.***) =** $\dfrac{6.4}{\text{Specific heat (in cals.)}}$

This law is applicable to solid elements only except Be, B, C and Si because their specific heat is variable with temperature.

Atomic mass = Equivalent mass $\times$ **Valency**

$$\text{Valency} = \frac{\text{Approximate atomic mass}}{\text{Equivalent mass}}$$

(ii) Vapour density method : It is suitable for elements whose chlorides are volatile.

Valency of the element = $\dfrac{\text{Molecular mass of chloride}}{\text{Equivalent mass of chloride}}$

$= \dfrac{2 \times \text{Vapour density of chloride}}{\text{Equivalent mass of metal} + 35.5}$

Atomic mass = Equivalent mass of metal $\times$ **Valency**

(iii) Specific heat method : It is suitable only for gases. The two types of specific heats of gases are C_P (at constant pressure) and C_v (at constant volume). *Their ratio is known as γ whose value is constant (1.66 for monoatomic, 1.40 for diatomic and 1.33 for triatomic gases).*

$$\text{Atomic mass of a gaseous element} = \frac{\text{Molecular mass}}{\text{Atomicity}}$$

(iv) Isomorphism method : It is based on law of isomorphism which states that compounds having identical crystal structure have similar constitution and chemical formulae.

Example : K_2SO_4, K_2CrO_4 and K_2SeO_4

(valency of **S, Cr, Se** = 6),

$ZnSO_4.7H_2O$, $MgSO_4.7H_2O$, $FeSO_4.7H_2O$

(valency of **Zn, Mg, Fe** = 2).

(2) Molecular mass : Molecular mass of a molecule, of an element or a compound may be defined as a *number which indicates how many times heavier is a molecule of that element or compound as compared with* $\frac{1}{12}$ *of the mass of an atom of carbon*–12. Molecular mass is a ratio and hence has no units. It is expressed in *a.m.u.*

$$\text{Molecular mass} = \frac{\text{Mass of one molecule of the substance}}{1/12 \times \text{Mass of one atom of C - 12}}$$

Actual mass of one molecule = Mol. mass $\times 1.66 \times 10^{-24}$ *gm.*

Molecular mass of a substances is the additive property and can be calculated by adding the atomic masses present in one molecule.

Gram molecular mass (GMM) and Gram molar volume : Molecular mass of an element or compound when expressed in *gm.* is called its gram molecular mass, gram molecule or mole molecule.

Number of *gm* molecules = $\frac{\text{Mass of substances}}{\text{GMM}}$

Mass of substances in *gm* = No. of *gm.* molecules $\times$ **GMM**

Volume occupied by one mole of any gas at STP is called **Gram molar volume**. The value of gram molar volume is 22.4 *litres*. *Volume of 1 mole of any gas at STP = 22.4 litres*

Expression for mass and density

Mass of $11.2L$ **of any gas at STP = V.D. of that gas in *gm.***

Density of a gas at NTP = $\frac{\text{Mol.mass in } gm.}{22400 \, ml}$

Important generalisations

Number of atoms in a substance

= Number of GMM $\times 6.02 \times 10^{23} \times$ **Atomicity**

Number of electrons in given substance

= Number of GMM $\times 6.02 \times 10^{23} \times$ **Number of electrons**

Methods of determination of molecular mass

Following methods are used to determine molecular mass,

(i) Diffusion method (For gases) : The ratio of rates of diffusion of two gases is inversely proportional to the square root of their molecular masses.

$$\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$$

(ii) Vapour density method (For gases only) : Mass of a fixed volume of the vapour is compared with the mass of the same volume of hydrogen under same conditions. The ratio of these masses is called **Vapour density** or **Relative density**.

$$\text{Molecular mass} = 2 \times \text{Vapour density}$$

(iii) Victor Meyer method (For volatile liquids or solids)

It is based on Dalton's law of partial pressure and Avogadro's hypothesis (gram molar volume).

22400 *ml* of vapours of a substance = Molecular mass of that substance

(iv) Colligative property method (For non-volatile solids)

Discussed in colligative properties of solutions.

Average atomic mass and molecular mass

$\overline{A}$ **(Average atomic mass) =** $\dfrac{\sum A_i X_i}{\sum X_{\text{total}}}$

$\overline{M}$ **(Average molecular mass) =** $\dfrac{\sum M_i X_i}{\sum X_{\text{total}}}$

Where $A_1, A_2, A_3 .....$ are atomic mass of species 1, 2, 3,.... etc. with % ratio as $X_1, X_2, X_3 .....$ etc. Similar terms are for molecular masses.

(3) Equivalent mass : The number of parts by mass of a substance that combines with or displaces 1.008 parts by mass of hydrogen or 8.0 parts of oxygen or 35.5 parts of chlorine is called its *equivalent mass* (EM). On the other hand *quantity of a substance in grams numerically equal to its equivalent mass is called its gram equivalent mass (GEM) or gram equivalent*.

$$\text{Number of GEM} = \frac{\text{Mass of the substance in grams}}{\text{GEM of the substance}}$$

Expressions for equivalent mass (EM)

(i) EM of an element $= \dfrac{\text{Atomic mass}}{\text{Valency}}$

(ii) EM of an acid $= \dfrac{\text{Molecular mass}}{\text{Basicity}}$

(*Basicity of acid is the number of replaceable hydrogen atoms in one molecule of the acid*).

(iii) EM of a base $= \dfrac{\text{Molecular mass}}{\text{Acidity}}$

(*Acidity of a base is the number of replaceable−OH groups in one molecule of the base*).

(iv) EM of a salt $= \dfrac{\text{Formula mass}}{\text{Total positive or negative charge}}$

(v) EM of an oxidising agent

$$= \frac{\text{Formula mass}}{\text{Number of electrons gained per molecule or Total change in O.N.}}$$

Equivalent mass of common oxidising agent changes with the medium of the reaction.

Methods of determination of equivalent mass

(i) Hydrogen displacement method : The mass of metal which displaces 11200 *ml* of hydrogen at NTP from an acid, alkali or alcohol is the equivalent mass of the metal.

(a) Equivalent mass of metal

$$= \frac{\text{Mass of metal}}{\text{Mass of H}_2 \text{ displaced}} \times 1.008 = \frac{W}{M} \times 1.008 \, g$$

(b) Equivalent mass of metal

$$= \frac{\text{Mass of metal}}{\text{Vol.(ml) of H}_2 \text{ displaced at STP}} \times 11200 = \frac{W}{V} \times 11200$$

This method is useful for metals which can displace hydrogen from acids or can combine with hydrogen (Mg, Zn, Na, Ca etc.)

(ii) Oxide formation method : The mass of the element which combines with 8 grams of oxygen is the equivalent mass of the element.

(a) Equivalent mass of metal $= \dfrac{\text{Mass of metal}}{\text{Mass of oxygen}} \times 8$

(b) Equivalent mass of metal $= \dfrac{\text{Mass of metal}}{\text{Vol. of } O_2 \text{ at S.T.P. in ml}} \times 5600$

(iii) Chloride formation method : The mass of an element which reacts with 35.5 *gm.* of chlorine is the equivalent mass of that element.

(a) Equivalent mass of metal $= \dfrac{\text{Mass of metal}}{\text{Mass of chlorine}} \times 35.5$

(b) Equivalent mass of metal $= \dfrac{\text{Mass of metal}}{\text{Vol. of } Cl_2 \text{ (in ml.) at STP}} \times 11200$

(iv) Neutralisation method : (For acids and bases).

$$\text{Equivalent mass of acid (or base)} = \frac{W}{V \times N}$$

Where ,

W = Mass of acid or base in *gm.,*

V = Vol. of base or acid in litre required for neutralisation

N is Normality of base or acid

(v) Metal displacement method : *It is based on the fact that one gm. equivalent of a more electropositive metal displaces one gm equivalent of a less electropositive metal from its salt solution.*

$$\frac{\text{Mass of metal added}}{\text{Mass of metal displaced}} = \frac{\text{Eq. mass of metal added}}{\text{Eq. mass of metal displaced}}$$

$$\frac{W_1}{W_2} = \frac{E_1}{E_2}$$

(vi) Electrolytic method : The quantity of substance that reacts at electrode when 1 faraday of electricity is passed is equal to its gram equivalent mass.

Gram equivalent mass = Electrochemical equivalent × 96500

The ratio of masses of two metals deposited by the same quantity of electricity will be in the ratio of their equivalent masses.

$$\frac{W_1}{W_2} = \frac{E_1}{E_2}$$

(vii) Double decomposition method

$$AB + CD \longrightarrow AD \downarrow + CB$$

$$\frac{\text{Mass of compound } AB}{\text{Mass of compound } AD} = \frac{\text{Eq. mass of } A + \text{Eq. mass of } B}{\text{Eq. mass of } A + \text{Eq. mass of } D}$$

or $\dfrac{\text{Mass of salt taken } (W_1)}{\text{Mass of ppt. obtained } (W_2)} = \dfrac{\text{Eq. mass of salt} (E_1)}{\text{Eq. mass of salt in ppt.} (E_2)}$

(viii) Conversion method : When one compound of a metal is converted to another compound of the same metal, then

$$\frac{\text{Mass of compound I} (W_1)}{\text{Mass of compound II} (W_2)} = \frac{E + \text{Eq. mass of radical I}}{E + \text{Eq. mass of radical II}}$$

(*E* = Eq. mass of the metal)

(ix) Volatile chloride method

Valency of metal $= \dfrac{2 \times \text{V.D. of Chloride}}{\text{Eq. mass of metal chloride}} = \dfrac{2 \times \text{V.D.}}{E + 35.5}$

$$\therefore \; \boldsymbol{E} = \frac{2 \times \text{V.D. of Chloride}}{\text{Valency}} - 35.5$$

(x) Silver salt method (For organic acids)

Equivalent Mass of acid = $\dfrac{108 \times \text{Mass of silver salt}}{\text{Mass of } Ag \text{ metal}} - 107$

Molecular mass of acid = Equivalent mass of acid × Basicity

The mole concept

One mole of any substance contains a fixed number (6.022×10^{23}) of any type of particles (atoms or molecules or ions) and has a mass equal to the atomic or molecular weight, in grams. Thus it is correct to refer to a mole of helium, a mole of electrons, or a mole of Na^+, meaning respectively *Avogadro's number* of atoms, electrons or ions.

$$\therefore \text{ Number of moles } = \frac{\text{Weight (grams)}}{\text{Weight of one mole (g/mole)}}$$

$$= \frac{\text{Weight}}{\text{Atomic or molecular weight}}$$

Percentage composition & Molecular formula

(1) Percentage composition of a compound

Percentage composition of the compound is the relative mass of each of the constituent element in 100 parts of it. If the molecular mass of a compound is *M* and *B* is the mass of an element in the molecule, then

$$\text{Percentage of element } = \frac{\text{Mass of element}}{\text{Molecular mass}} \times 100 = \frac{X}{M} \times 100$$

(2) Determination of empirical formula : The empirical formula of a molecule is determined using the % of elements present in it. Following method is adopted.

Element	%	Relative no. of atoms = %/at. wt.	Simplest Ratio	Empirical Formula

Relative no. of atoms : Divide the percentage of each element present in compound by its at. weight. This gives the relative no. of atoms of element in molecule.

Simplest ratio : Find out lowest value of relative no. of atoms and divide each value of relative no. of atoms by this value to estimate simplest ratio of elements.

It the simplest ratio obtained are not complete integers, multiply them by a common factor to get integer values of simplest ratio.

Empirical formula : Write all constituent atoms with their respective no. of atoms derived in simplest ratio. This gives empirical formula of compound.

Molecular formula : Molecular formula $= n \times$ empirical formula where $'n'$ is the whole no. obtained by

$$n = \frac{\text{molecular weight of compound}}{\text{empirical formula weight of compound}}$$

(iii) *Iodiometric titrations* : This is a simple titration involving free iodine. This involves the titration of iodine solution with known sodium thiosulphate solution whose normality is N. Let the volume of sodium thiosulphate is $V\,ml$.

$$I_2 + 2Na_2S_2O_3 \rightarrow 2NaI + Na_2S_4O_6$$
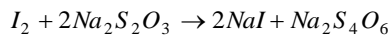
$n = 2$, $n = 1$

Equivalents of I_2 = Equivalent of $Na_2S_2O_3$

$\therefore$ **Equivalents of $I_2 = N \times V \times 10^{-3}$**

Moles of $I_2 = \dfrac{N \times V \times 10^{-3}}{2}$

Mass of free I_2 in the solution $= \left[\dfrac{N \times V \times 10^{-3}}{2} \times 254\right] g$.

(iv) *Iodometric titrations* : This is an indirect method of estimation of iodine. An oxidising agent is made to react with excess of solid KI. The oxidising agent oxidises I^- to I_2. This iodine is then made to react with $Na_2S_2O_3$ solution.

Oxidising Agent

$$(A) + KI \rightarrow I_2 \xrightarrow{2Na_2S_2O_3} 2NaI + Na_2S_4O_6$$
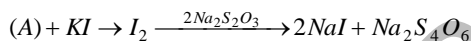
Let the normality of $Na_2S_2O_3$ solution is N and the volume of thiosulphate consumed to $V\,ml$.

Equivalent of A = Equivalent of I_2 = Equivalents of $Na_2S_2O_3$

Equivalents of I_2 liberated from $KI = N \times V \times 10^{-3}$

Moles of I_2 liberated from $KI = \dfrac{N \times V \times 10^{-3}}{2}$

Mass of I_2 liberated from $KI = \left[\dfrac{N \times V \times 10^{-3}}{2} \times 254\right] g$.

Formula used in solving numerical problems on volumetric analysis

(1) Strength of solution = Amount of substance in $g\ litre^{-1}$

(2) Strength of solution = Amount of substance in $g\ moles\ litre^{-1}$

(3) Strength of solution = Normality $\times$ Eq. wt. of the solute

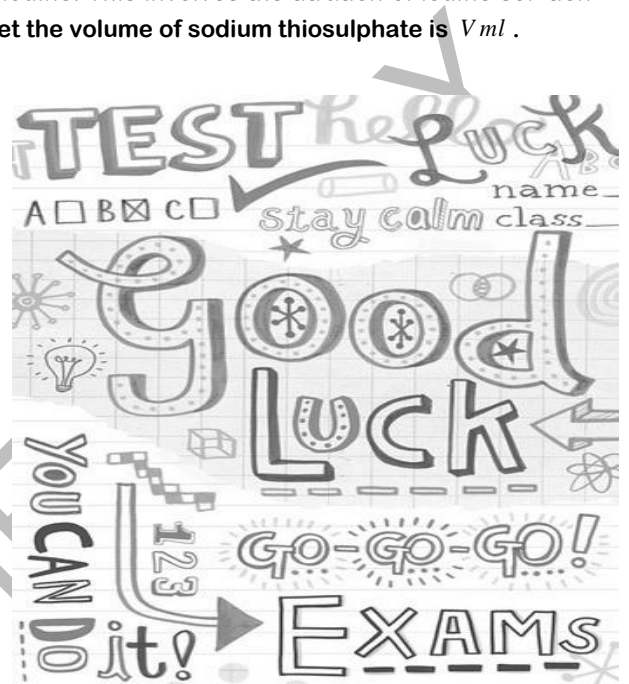= molarity $\times$ Mol. wt. of solute

(4) Molarity $= \dfrac{\text{Moles of solute}}{\text{Volume in litre}}$

(5) Number of moles $= \dfrac{\text{Wt. in } gm}{\text{Mol.wt.}} = M \times V_{(inl)}$

$= \dfrac{\text{Volume in litres}}{22.4}$ **at NTP (only for gases)**

(6) Number of millimoles $= \dfrac{\text{Wt. in } gm \times 1000}{\text{mol. wt.}}$

$=$ Molarity$\times$ Volume in ml.

(7) Number of equivalents

$$= \frac{\text{Wt.in } gm}{\text{Eq.wt.}} = x \times \text{No. of moles} \times \text{Normality} \times \text{Volume in litre}$$

(8) Number of milliequivalents (meq.)

$$= \frac{\text{Wt.in } gm \times 1000}{\text{Eq.wt.}} = \text{normality} \times \text{Volume in } ml.$$

(9) Normality $= x \times \text{No. of millimoles}$

$$= x \times \text{Molarity} = \frac{\text{Strength in } gm\, litre^{-1}}{\text{Eq.wt.}}$$

where $x = \dfrac{\text{Mol.wt.}}{\text{Eq.wt.}}$, x **= valency or change in oxi. Number.**

(10) Normality formula, $N_1 V_1 = N_2 V_2$

(11) % by weight $= \dfrac{\text{Wt.of solvent}}{\text{Wt.of solution}} \times 100$

(12) % by volume $= \dfrac{\text{Wt.of solvent}}{\text{Vol.of solution}} \times 100$

(13) % by strength $= \dfrac{\text{Vol.of solvent}}{\text{Vol.of solution}} \times 100$

(14) Specific gravity

$$= \frac{\text{Wt.of solution}}{\text{Vol.of solution}} = \text{Wt.of 1 } ml. \text{ of solution}$$

(15) Formality $= \dfrac{\text{Wt.of ionic solute}}{\text{Formula Wt.of solute} \times V_{inl}}$

(16) Mol. Wt. = V.D $\times$ **2 (For gases only)**