Mole Concept

BY – Nitesh Jain(IIT Bombay)

Mole Concept

Mole is defined as a ‘defined number of particles. This definite number is called Avogadro constant $(N_A) = 6.028 \times 10^{23}$

NOTE: just as, $1\ dozen \Rightarrow 12\ objects$

$1\ score \Rightarrow 20\ objects$

$1\ mole \Rightarrow 6.028 \times 10^{23}$ atoms/ molecules/ions etc……

A gram- molecule → termed as a **mole of molecules**

A gram- atom → termed as **a mole of atoms**

Number of moles of molecules ⇨ $= \dfrac{weight\ in\ g}{molecular\ weight}$

Number of moles of atoms ⇨ $= \dfrac{weight\ in\ g}{atomic\ weight}$

Rules of Mole Concept

Number of moles of atoms/molecules/ions/electrons

$$= \frac{\boldsymbol{no.of\ atoms\ or\ molecules\ or\ ions\ or\ electrons}}{\boldsymbol{Avogadro\ constant\ (N_A)}}$$

$$where\ N_A = 6.028 \times 10^{23}$$

For a compound M_xN_y

$$x\ moles\ of\ M = y\ moles\ of\ N$$

Atomic and Molecular Mass

Atomic mass of element $= A$

$$= \frac{mass\ of\ one\ \ atom\ of\ the\ element}{mass\ of\ one\ atom\ of\ hydrogen}$$

Atomic mass is a number which tells how many times the element is heavier w.r.t mass of hydrogen

In modern concept, the reference is changed from hydrogen to $\frac{\mathbf{1}}{\mathbf{12}}^{\boldsymbol{th}}$ part of mass of one atom of $\boldsymbol{carbon - 12}$, which is referred to as $\boldsymbol{1\ amu}$ $(\boldsymbol{1amu = 1.66 \times 10^{-27}\ kg})$

Example: atomic mass of oxygen= $16\ amu$

& actual mass of oxygen $= 16 \times 1.66 \times 10^{-27}\ kg$

Similarly

$$\boldsymbol{molecular\ mass = \frac{mass\ of\ one\ molecule\ of\ the\ substance}{1\ amu}}$$

Standard Molar Volume ($S.M.V$)

$S.M.V$ is the volume occupied by $\mathbf{1}\ \boldsymbol{mole}$ of any gas at NTP

Value of $\boldsymbol{S.M.V = 22.4\ litres}$

NOTE: NTP is condition when $T = 0°C$ and $P = 1\ atm$

$$\therefore\ number\ of\ moles\ of\ gas\ = \frac{\boldsymbol{Volume\ at\ NTP}}{\boldsymbol{standard\ molar\ volume}}$$

Principle of Atom Conservation ($POAC$)

Conservation of atoms and it means moles of atoms shall also be conserved

NOTE: This concept is the basis of mole-concept

It helps in

- Stoichiometric calculations
- Balancing of chemical equations

Application of $POAC$

Example: $KClO_3(s) \rightarrow KCl(s) + O_2(g)$

Applying $S.M.V$ for K atoms

$$moles\ of\ K\ atoms\ in\ reactant\ =\ moles\ of\ K\ atoms\ in\ products$$

Since 1 mole of $KClO_3$ contains 1 atom of K;

1 $mole$ of $KClO_3 =$ 1 $mole\ of\ K$

1 $mole\ of\ KCl =$ 1 $mole\ of\ K$

$$\therefore moles\ of\ KClO_3 = moles\ of\ KCl$$

$$\frac{weight\ of\ KClO_3}{molecular\ weight\ of\ KClO_3} = \frac{weight\ of\ KCl}{molecular\ weight\ of\ KCl}$$

NOTE: Hence, using **POAC**, weight relationship between $KClO_3$ and KCl can be obtained, which is useful for stoichiometric calculations

There is no need to balance the relation and directly $POAC$ can be applied

Homogeneous Gaseous Reactions

$$H_2(g) + I_2(g) \rightarrow 2HI(g)$$

	1 molecule	1 molecule	2 molecules
or			
	1 mole	1 mole	2 moles

Using Avogadro's principle

1 volume	1 volume	2 volumes	At T & P constant
1 pressure	1 pressure	2 pressures	At T & V constant

Avogadro's principle states that under same conditions of T&P, equal conditions of gases contain same number of molecules

Hence, relative volumes of each reactant and product can be obtained under same T & P conditions

NOTE: Relative weights of reactant and product can be found out using principle of Atom Conservation

Law of Chemical Equivalence

The law states that one equivalent of an element combines with one equivalent of the other

An equivalent of a substance is that amount of substance which combines with $1\ mole$ of hydrogen atoms or $8\ mole\ Oxygen\ atoms$

$atomic\ weight\ of\ Br$

Example: equivalent weight of $Br = 1 \times 79.9 = 79.9\ (Since\ it\ forms\ HBr)$

equivalent mass is pure number. When expresses in grams, it is called gram equivalent mass

Also, it may vary for a substance with change in velocity

For $CuO = 31.75$

For $Cu_2O = 63.5$

Relation between Atomic Mass ,Equivalent Mass and Velocity

Let an element X combine with H to form $\boldsymbol{XH_n}$

$\downarrow$

Valance of X

$$\therefore \boldsymbol{Equivalent\ mass} = \frac{\boldsymbol{Atomic\ mass}}{\boldsymbol{n}}$$

From definition of equivalent mass

1 $part$ by mass of hydrogen combines with $\left(\frac{\boldsymbol{atomic\ mass\ of\ X}}{\boldsymbol{n}}\right)$ $parts\ of\ X$

$$\boxed{\boldsymbol{Atomic\ mass} = \boldsymbol{valency} \times \boldsymbol{Equivalent\ mass}}$$

Types of Formulae for Compounds

Empirical Formula → It represents the simplest relative whole number ratio of atoms of each element present in the molecule of compound

Example: for $C_6H_{12}0_6$, empirical formula is CH_20

Molecular Formula → Represents actual number of atoms of each element present in the compound

$$= \boldsymbol{n \times Empirical\ formula}$$

$$n = \frac{molecular\ formula\ mass}{emperical\ formula\ mass}$$

Example: $n = 6$ for $C_6H_{12}0_6$

Determination of Empirical Formula

1. Find percentage combination of compounds

It is the relative mass of each of constituent element of compound in 100 parts of the compound

2. Percentage of each element is divided by its atomic mass

This value gives atomic ratio of elements present

3. Simplest ratio of atoms of elements is found by dividing atomic ratio of each element by minimum value of atomic ratio

4. Convert simplest ratio to whole number, If it is fractional

Empirical formula = symbols of elements present written side by side with their respective whole number ratio

Concept of Limiting Reagent

In the reactions, where more than one reactant is involved, the reactant which is completely consumed is **called limiting reagent**

Example: $A \quad + \quad 2B \longrightarrow 4C$

At time$= 0$ $\quad 5\,moles \quad 12\,moles$

A gives 20 moles of C (if A is limiting reagent)

B gives 24 moles of C (if B is limiting reagent)

NOTE: Limiting reagent decides that what quantity of products will be formed

Reactant producing least number of moles of the product is the limiting reactant ($\therefore A$ in above example)

Dulong and Petit's Law

Atomic mass $\times$ specific heat $= 6.4$

NOTE: Applicable for a solid element

Also, the unit of specific heat in formula above, must be in cal/g unit

Equivalent Weight for Electrolytes

$$\boldsymbol{E_{(electrolytes)} = E_{cations} + E_{anions}}$$

NOTE: For electrolytes, it is assumed that it undergoes complete ionization

Example

$$E_{NaCl} = E_{Na^+} + E_{Cl^-}$$

Vapor density

$$Absolute\ density = \frac{mass}{volume}$$

$$Relative\ density = \frac{absolute\ density}{density\ of\ pure\ water\ at\ 4℃}$$

$$\boldsymbol{vapor\ density} = \frac{mass\ of\ vapor\ of\ substance\ per\ mLat\ NTP}{mass\ of\ hydrogen\ per\ mL\ \ at\ NTP}$$

Since both, vapor of substance and hydrogen occupy equal volume at NTP

Or $$\boldsymbol{vapor\ density} = \frac{\boldsymbol{molecular\ mass\ of\ vapor}}{\boldsymbol{molecular\ mass\ of\ hydrogen}}$$

$\frac{molecular\ mass}{2}$ (since, molecular mass of $H_2 = 2$)

From vapor density calculation, molecular mass of a compound can be found out

Law of Chemical Equivalence in a Chemical Reaction

$$aA + bB \rightarrow mM + nN$$

The equivalent weight (in grams) of a compound taking part in a reaction	=	weight of compound which combines with 1 equivalent of another compound

$$eq.of\ A = eq.of\ B = eq.of\ M = eq.of\ N$$

$$\boldsymbol{\frac{wt.of\ A}{E_A} = \frac{wt.of\ B}{E_B} = \frac{wt.of\ M}{E_M} = \frac{wt.of\ N}{E_N}}$$

Where E stands for equivalent weight

Equivalent Weight for Acids and Bases

For Acids,

$$E = \frac{\boldsymbol{Molecular\ weight\ of\ acids(i.e.weight\ of\ 1\ mole)}}{\boldsymbol{basicity\ of\ Acid}}$$

i.e. mole of H^+ furnished

For Bases,

$$E = \frac{\boldsymbol{Molecular\ weight\ of\ base}}{\boldsymbol{acidity\ of\ Base}}$$

i.e. moles of OH^- furnished or moles of H^+ accepted

Equivalent Weight for Salts

$$E = \frac{molecular\ weight\ of\ salt}{moles\ of\ H\ equivalent\ to\ total\ number\ of\ cations\ or\ anions}$$

or

$$E = \frac{molecular\ weight\ of\ salt}{moles\ of\ metal\ atoms\ \times valency\ of\ metal}$$

Example 1: For $CaSO_4$, $E = \frac{M}{2}$

$$for\ oxides: E = \frac{molecular\ weight\ of\ oxide}{moles\ of\ element\ atoms\ \times valency\ of\ element}$$

Example 2: For Al_2O_3, $E = \frac{M}{6}$

NOTE: Equivalent weight of a compound taking part in a reaction should be determined from chemical equation, as it depends on stoichiometry of the reaction

Equivalent Weight in Redox Reactions

In redox reactions, equivalent weight $(E) = \frac{Mol.wt\ of\ oxidising\ or\ reducing\ agent}{change\ in\ oxidation\ no.per\ mole\ of\ Redox\ agent}$

Example: when $MnO_4{}^-$ is reduced to Mn^{2+} → oxidation number changes from $+7\ to + 2$

$$\boldsymbol{E = \frac{M_{MnO_4^-}}{5}}$$

NOTE: Oxidation number of ions taking part in the precipitation reaction

does not change. For them $\boldsymbol{E = \frac{weight\ of\ 1\ mole\ of\ ions}{mole(s) of\ charge\ on\ ion}}$

Example: $E_{Fe^{2+}} = \frac{56}{2} = 28$ AND $E_{Fe^{3+}} = \frac{56}{3} = 18.67$

Molecular weight of a Compound

It is defined as the weight of a molecule of the compound relative to a carbon atom

NOTE: Molecular weight in grams is the weight of 1 mole of molecule

$$molecular\ weight = \frac{weight\ of\ molecules\ in\ grams}{number\ of\ moles\ of\ molecules}$$

It is measured in $\boldsymbol{atomic\ mass\ unit}\ (\boldsymbol{amu})$

Atomic Weight of an Element

Atomic weight is the average weight of atoms of the element relative to a carbon atom, taken as exactly 12

NOTE: It is the weight of one mole of atoms

formula

$$\therefore \text{mathematically,} \qquad \text{atomic weight} = \frac{weight\ of\ atoms\ in\ grams}{number\ of\ moles\ of\ atoms}$$

It is measured in atomic mass unit (amu)

1 amu is defined as $\frac{1}{12}^{th}$ of the mass of ^{12}C isotope

1 amu is also called **one Dalton**