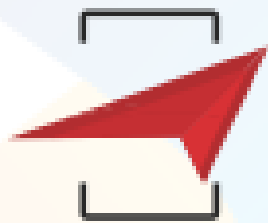


Mole Concept



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An Initiative by अमर उजाला

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Mole Concept

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

NOTE: just as, 1 dozen \Rightarrow 12 objects
1 score \Rightarrow 20 objects
1 mole \Rightarrow 6.028×10^{23} atoms/ molecules/ions etc.....

A gram- molecule \rightarrow termed as a **mole of molecules**

A gram- atom \rightarrow termed as **a mole of atoms**

Number of moles of molecules \rightarrow

$$= \frac{\text{weight in g}}{\text{molecular weight}}$$

Number of moles of atoms \rightarrow

$$= \frac{\text{weight in g}}{\text{atomic weight}}$$

Rules of Mole Concept

Number of moles of atoms/molecules/ions/electrons

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

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

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

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 $\frac{1}{12}^{th}$ part of mass of one atom of *carbon* – 12, which is referred to as **1 amu** ($1\text{amu} = 1.66 \times 10^{-27} \text{ kg}$)

Example: atomic mass of oxygen = 16 amu

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

Similarly

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

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

S.M.V is the volume occupied by **1 mole** of any gas at NTP



Value of *S.M.V* = **22.4 litres**



NOTE: *NTP* is condition when $T = 0^{\circ}\text{C}$ and $P = 1 \text{ atm}$

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

Principle of Atom Conservation (*POAC*)

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

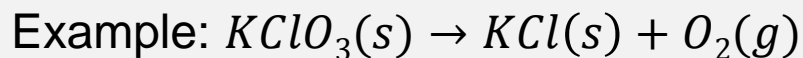
NOTE: This concept is the basis of mole-concept

It helps in

Stoichiometric calculations

Balancing of chemical equations

Application of *POAC*



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*



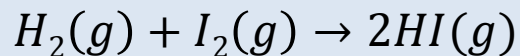
\therefore moles of $KClO_3 =$ moles of KCl

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

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



or

1 molecule 1 molecule 2 molecules

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 expressed in grams, it is called gram equivalent mass

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

For *CuO* = 31.75

For *Cu₂O* = 63.5

Relation between Atomic Mass ,Equivalent Mass and Velocity

Let an element X combine with H to form XH_n

\downarrow
Valance of X

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

From definition of equivalent mass

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

$$\text{Atomic mass} = \text{valency} \times \text{Equivalent mass}$$

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 $C_6H_{12}O_6$, empirical formula is CH_2O

Molecular Formula



Represents actual number of atoms of each element present in the compound

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

$$n = \frac{\textit{molecular formula mass}}{\textit{empirical formula mass}}$$

Example: $n = 6$ for $C_6H_{12}O_6$

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**



At time = 0 5 moles 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 reagent (\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

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

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

Example

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

Vapor density

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

$$\text{Relative density} = \frac{\text{absolute density}}{\text{density of pure water at } 4^{\circ}\text{C}}$$

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

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

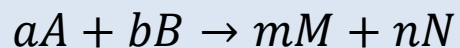
Or

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


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

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

Law of Chemical Equivalence in a Chemical Reaction



The equivalent weight (in grams) of a compound taking part in a reaction



weight of compound which combines with 1 equivalent of another compound

$$eq. \text{ of } A = eq. \text{ of } B = eq. \text{ of } M = eq. \text{ of } N$$

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

Where E stands for equivalent weight

Equivalent Weight for Acids and Bases

For Acids,

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

i.e. mole of H^+ furnished

For Bases,

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

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

Equivalent Weight for Salts

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

or

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

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

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

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) = $\frac{\text{Mol. wt of oxidising or reducing agent}}{\text{change in oxidation no. per mole of Redox agent}}$

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

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

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

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

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



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

It is measured in *atomic mass unit (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

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

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**