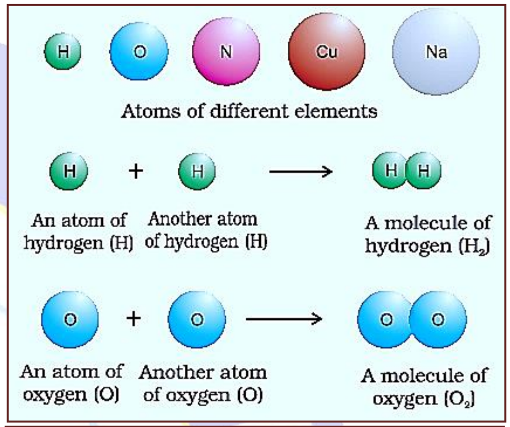
Chapter 01
Some Basic Concepts of Chemistry

CONTENT

- **Classification of universe**
- **Measurement of physical properties and SI units system**
- **Mole Concept**
- **Percentage composition, Empirical & Molecular formula**
- **Chemical equation**
- **Stoichiometry**
- **Limiting Reagent**
- **Equivalent weight**
- **Concentration terms**
- **Dalton's atomic theory and laws of chemical combination.**

INTRODUCTION :



Arrangement of atoms and molecules

- Chemistry deals with the composition, structure and properties of matter.
- These aspects can be best described and understood in terms of basic constituents of matter: **atoms** and **molecules.**
- That is why chemistry is called the science of atoms and molecules. Can we see, weight and perceive these entities?
- Is it possible to count the number of atoms and molecules in a given mass of matter and have a quantitative relationship between the mass and number of these particles (atoms and molecules)? We will like to answer some of these questions in this Unit.
- We would further describe how physical properties of matter can be quantitatively described using numerical values with suitable units.

MATTER

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CLASSIFICATION

- **PHYSICAL**
  - **Solid**: definite shape & volume
  - **Liquid**: definite volume but no definite shape
  - **Gas**: no definite shape & volume
- **CHEMICAL**
  - **Pure Substance**: fixed ratio of masses of constituents
    - **Elements**: Consists of only one type of atoms
    - **Compounds**: Made of two or more atoms of different elements
  - **Mixture**
    - **Homogeneous**: same composition throughout & component are indistinguishable e.g. Gases, Alloy, Atmospheric air
    - **Heterogeneous**: not uniform e.g. Sand & water

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CLASSIFICATION OF UNIVERSE:

(A) Matter (B) Energy

(A) Matter: The thing which occupy space and having mass which can be felt by our five senses is called matter.

Matter is further classified into two categories:

(I) Physical classification

(II) Chemical classification

PHYSICAL CLASSIFICATION

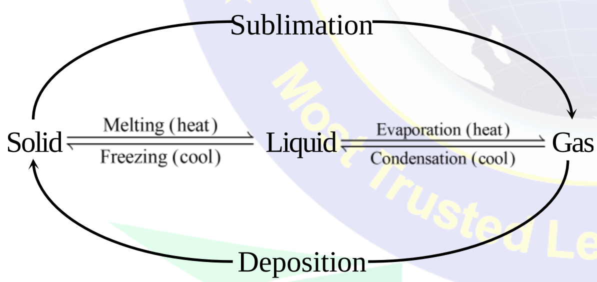
It is based on physical state under ordinary conditions of temperature and pressure, so on the basis of the nature of forces matter can be classified into the following three ways :

(A) Solid (B) Liquid (C) Gas

(A) Solid: A substance is said to be solid if it possesses a definite volume and a definite shape.
e.g. sugar, iron, gold, wood etc.

(B) Liquid: A substance is said to be liquid if it possesses a definite volume but not definite shape. Liquid takes up the shape of the vessel in which it is kept.
e.g. water, milk, oil, mercury, alcohol etc.

(C) Gas: A substance is said to be gas if it neither possesses a definite volume nor a definite shape. This is because they fill up the whole vessel in which they are put.
e.g. hydrogen(H_2), Oxygen(O_2), carbon dioxide(CO_2) etc.



CHEMICAL CLASSIFICATION

It may be classified into two types:

(A) Pure Substance

A material containing only one type of substance. Pure Substance can not be separated into simpler substance by physical method.

Pure substance is classified into two types :

(a) Element

(b) Compound

e.g.: Element = Na, Mg, Ca .... etc.

Compound = HCI, H_2O, CO_2, HNO_3 ....etc.

(i) Element: The pure substance containing only one kind of atoms.

It is classified into 3 types (depend on physical and chemical property)

(i) Metal $\rightarrow$ Zn, Cu, Hg, Ac, Sn, Pb etc.

(ii) Non-metal $\rightarrow$ N_2, O_2, Cl_2, Br_2, F_2, P_4, S_8 etc.

(iii) Metalloids $\rightarrow$ B, Si, As, Te etc.

(ii) Compound: It is defined as pure substance containing more than one kind of elements or atoms which are combined together in a fixed proportion by weight and which can be decomposed into simpler substance by the suitable chemical method. The properties of a compound are completely different from those of its constituent element.

e.g. HCI, H_2O, H_2SO_4, $HClO_4$, HNO_3 etc.

(B) Mixture

A material which contain more than one type of substances and which are mixed in any ratio by weight is called as mixture. The property of the mixture is the property of its components. The mixture can be separated by simple physical method.

Mixture is classified into two types :

(i) Homogeneous mixture: The mixture, in which all the components are present **uniformly** is called as homogeneous mixture. Components of mixture are present in single phase.

e.g. Water + Salt, Water + Sugar, Water + alcohol.

(ii) Heterogenous mixture: The mixture in which all the components are present **non-uniform**

e.g. Water + Sand, Water + Oil, blood, petrol etc.

MEASUREMENT OF PHYSICAL PROPERTIES AND SI UNITS SYSTEM

- Chemists describe the behaviour of chemical substances on the basis of physical and chemical properties.
- **The measurements of chemical properties involve chemical reactions**, whereas the measurement of physical properties does not involve any chemical reactions.
- The common physical properties are **mass, length, time, volumes, temperature, density, etc., among these mass, lengths and time** are **fundamental physical quantities**.

(1) Mass tells us about the quantity of matter. Mass is measured with the help of analytical balance.

(2) Size of the object is measured in terms of **length, area and volume**. Length refers to one dimension, area to two dimensions and volume to three dimensions of space.

(3) **Time** helps us to know how long it takes for a process to occur.

Seven basic units of measurement namely **mass length, time, temperature, electric current, luminous intensity and amount of substance** are taken as **fundamental basic units**. All the other units can be derived from them are called **derived units** like **area, volume, force, work, density, velocity, energy**, etc., are all **derived units**.

➢ **S.I. UNITS:** has seven basic units from which all other units are derived.

Physical quantity	Name of unit	Abbreviation
Mass	Kilogram	kg
Length	Meter	m
Temperature	Kelvin	K
Amount of substance	Mole	mol
Time	Second	s
Electric current	Ampere	A
Luminous intensity	Candela	cd

➢ **DERIVED UNITS: -** Units of different physical quantities can be derived from the seven basic units are called **derived units** because these are derived from the basic units.

For deriving these units, we can multiply or divide the symbols for units as if they are algebraic quantities

- **Area** = Length x Breadth = m x m = $\mathbf{m^2}$
- **Volume** = Length x Breadth x Height
  If units of length are m, then V = m x m x m = $\mathbf{m^3}$
- **Density** = $\frac{\text{Mass}}{\text{Volume}} = \frac{\mathbf{kg}}{\mathbf{m^3}}$
- **Velocity** = $\frac{\text{Distance}}{\text{Time}} = \frac{m}{s} = \mathbf{m.s^{-1}}$
- **Acceleration** = $\frac{\text{Velocity}}{\text{Time}} = \frac{m.s^{-1}}{s} = \mathbf{m.s^{-2}}$
- **Force** = Mass x acceleration = $\mathbf{kg.\ m.s^{-2}}$
- **Pressure** = $\frac{\text{Force}}{\text{Area}} = \frac{kg.m.s^{-2}}{m^2} = \mathbf{kg.m^{-1}.s^{-2}}$
- **Energy, work** = Force x distance = $kg.m.s^{-2}$ x m = $kg.m^2.s^{-2}$ = **Joule**
- **Electric charge** = current x time = A.s = **Coulomb**
- **Electric potential** = $\frac{\text{Energy}}{\text{Charge}} = \frac{kg.m^2.s^{-2}}{A.s}$ = Joule. $A^{-1}.s^{-1}$ = **Volt**

Quantity with Symbol	Unit (S.I.)	Symbol
Velocity, v	Metre per sec	ms^{-1}
Area, A	Square metre	m^2
Volume, V	Cubic metre	m^3
Density, ρ	Kilogram m^{-3}	kg m^{-3}
Energy, E	Joule (J)	kg m^2s^{-2}
Force, F	Newton (N)	kg ms^{-2}
Frequency, ν	Hertz	Cycle per sec (Hz)
Pressure, P	Pascal (Pa)	Nm^{-2}
Electrical charge	Coulomb (C)	A-s (ampere –sec)

➢ **PREFIXES:** The SI units of some of the physical quantities are either too small or too large. To change the order of magnitude, these are expressed by using prefixes before the names of the basic units.

Multiple	Prefix	Symbol	multiple	Prefix	Symbol
10^{24}	Yotta	Y	10^{-1}	deci	d
10^{21}	Zetta	Z	10^{-2}	centi	c
10^{18}	exa	E	10^{-3}	milli	m
10^{15}	peta	P	10^{-6}	micro	μ
10^{12}	tera	T	10^{-9}	nano	n
10^{9}	giga	G	10^{-12}	pico	p
10^{6}	mega	M	10^{-15}	femto	f
10^{3}	kilo	k	10^{-18}	atto	a
10^{2}	hecto	h	10^{-21}	zeto	z
10^{1}	deca	da	10^{-24}	yocto	y

UNITS AND DIMENSIONAL ANALYSIS

➢ **Conversion of Units:** It is used **to convert one set of units to another in calculations**. In order to use this method, we write the units with every number and carry the units through the calculations, treating them as algebraic quantities.

For interconversion of the units of time, we know that

1 min. = 60 sec.

$1 = \frac{60 \text{ sec}}{1 \text{ min}}$ or $1 = \frac{1 \text{ min}}{60 \text{ sec}}$

These equalities are called **unit conversion factor or conversion factor or simply unit factor**.

Other important Units

1 angstrom (Å) = 10^{-8} cm = 10^{-10} m = 10^{-1} nm = 10^{2} pm

1 inch = 2.45 cm or 1 cm = 0.394 inch

39.37 inch = 1 metre or 1 km = 0.621 mile

1 foot = 12 inch = 30.48 cm = 0.3048 m

1 kg = 2.20 pounds (lb) 1 g = 0.0353 ounce (oz)

1 pound (lb) = 453.6 g = 0.4536 kg

1 oz = 28.33 g = 0.02833 kg

1 metric ton = 1000 kg = 10^{6} g

1 atomic mass unit (amu) = 1.6605×10^{-24} g
= 1.6605×10^{-27} kg
= 1.492×10^{-3} erg
= 1.492×10^{-10} J
= 3.564×10^{11} cal
= 9.310×10^{8} eV
= 931.48 MeV

1 atmosphere (atm) =760 torr = 760 mm Hg = 76 cm Hg
= 1.01325×10^{5} Pa

1 calorie (cal) = 4.18400×10^{7} erg = 4.184 J
= 2.613×10^{19} eV

1 coulomb (coul) = 2.9979×10^{9} esu

1 electron volt (eV) = 1.602×10^{-12} erg
= 1.6021×10^{-19} J
= 3.827×10^{-20} cal
= 23.06 kcal mol^{-1}

1 erg (erg) = 10^{-7} J = 2.389×10^{-8} cal
= 6.242×10^{11} eV

1 dyne (dyne) = 10^{-5} N

1 joule (J) = 10^{7} erg = 0.2390 cal

1 litre (L) = 1000 cc
= 1000 mL = 1 dm^{3}
= 10^{-3} m^{3}

The conversion of units involves the following steps:

1. Determine, First of all, unit conversion factor, For **Ex:**

(a) 1 inch = 2.54 cm

Conversion factor: $\frac{1\,\text{inch}}{2.54\,\text{cm}}$ or $\frac{2.54\,\text{cm}}{1\,\text{inch}}$

(b) 1 lb = 454.0 g

Conversion factor: $\frac{1\,\text{lb}}{454.0\,\text{g}}$ or $\frac{454.0\,\text{g}}{1\,\text{lb}}$

(c) 1 calorie = 4.184 J

Conversion factor : $\frac{1\,\text{calorie}}{4.184\,\text{J}}$ or $\frac{4.184\,\text{J}}{1\,\text{calorie}}$

2. Multiply the given physical quantity with the appropriate unit conversion factor, (The appropriate conversion factor is so selected that it has the unit in the denominator which is to be converted). For Example, if x mL is to be converted into litre, the appropriate conversion factor is $\frac{0.001\,\text{L}}{1\,\text{mL}}$ and not $\frac{1\,\text{mL}}{0.001\,\text{L}}$. In case the unit conversion factor is not used correctly, the answer comes out with wrong units.

3. If the conversion involves more than one step, in each step, the conversion factor is used in such a way that the unit of the preceding factor cancel out. In calculation, units are always written along with the numbers and cancelled in the same manner as numbers.

Ex: To find the number the seconds in 5 min.

5min = 5 min $\frac{60\text{ sec}}{1\text{ min}}$ = 300 sec

Ex: To convert 0.74 Å into picometre.

1 Å = 10^{-10} m or 1=$\frac{10^{-10}\text{m}}{1\text{ Å}}$

0.74 Å = 0.74 Å x $\frac{10^{-10}\text{m}}{1\text{Å}}$= 0.74 x 10^{-10} m

1pm = 10^{-12} m or 1 = $\frac{1\text{ pm}}{10^{-12}\text{m}}$

$\therefore 0.74 \times 10^{-10}$m= 0.74×10^{-10}m $\times$ 10^{12}pm= 0.74×10^{2}pm= 74pm

- **Conversion of litre - atmosphere to joule**

1L = 10^{-3} m^{3} or 1 =$\frac{10^{-3}\text{m}^{3}}{1\text{ L}}$

1L atm = 1 L atm x $\frac{10^{-3}\text{m}^{3}}{1\text{ L}}$ = $10^{-3}$$m^{3}$atm

1atm = 101, 325 Pa or 1 = $\frac{101{,}325\text{ pa}}{1\text{ atm}}$

$10^{-3}$$m^{3}$atm=$10^{-3}$$m^{3}$atmx $\frac{101{,}325\text{ pa}}{1\text{ atm}}$ = 101, 325 Pa $\times 10^{-3}$ m^{3} = 101.325 Pa.m^{3}

But, Pa = $\frac{\text{N}}{\text{m}^{2}}$

$101.325 Pa.m^3 = 101.325.m^3 \frac{N}{m^2} = 101.325\ N.m$

$= 101.325\ J\ (\because 1\ J = N.m)$

UNCERTAINTY IN MEASUREMENT

➢ **Precision and Accuracy:**

- To express the results of different scientific measurements two terms accuracy and precision are commonly used.
- **Accuracy** is **a measure of the difference between the experimental value and the true value**.
  Small difference between the experimental value and the true value, larger is the accuracy.
  Accuracy expresses the correctness of measurement.
- **Precision** is **the difference between a measured value and the arithmetic mean value for a series of measurements**.
- Precision refers for the closeness of the set of values obtained from identical measurement of a quantity.
  Ex: Three students were asked to determine the mass of piece of metal where mass is known to be 0.520g. Data obtained by each student are recorded as follows.

Student. No.	Measurements (g)			Average(g)
Student A	0.521	0.515	0.509	0.515
Student B	0.516	0.515	0.514	0.515
Student C	0.521	0.520	0.520	0.520

- Data of student A are neither very precise nor accurate. The individual values differ widely from one another and the average value is not accurate.
- Student B was able to determine the mass more precisely. The values deviate, but little from one another. However, the average mass is still not accurate. The data for student C is both precise and accurate.

➢ **SCIENTIFIC NOTATION (OR) EXPONENTIAL NOTATION**

- In scientific notation, the large or small numbers are expressed in **N × 10ⁿ format** or a number between 1.000 and 9.999 multiplied or divided by 10, an appropriate no. of times.

Where, N is a no. between 1.000 and 9.999 and n is exponent.

Ex:

1) $138.42 = 1.3842\times10\times10 = 1.3842\times10^2$

2) $0.00013842 = 1.3842\times10^{-4}$

1. To transform a number **larger than 9.999..**. to scientific notation, the **decimal point there is only one non-zero digit before the decimal point. If the decimal point is moved x places to the left,** then exponent n = x.

Ex: $138.42 = 1.3842\times10^2$

$1395.2 = 1.3952\times10^3$

$21.654 = 2.1654\times10^1$

2. To transform a number **smaller than 1** to scientific notation, the **decimal point is moved to the right** until there is one nonzero digit before the decimal point. If the decimal point is moved 'y' places to the right, the exponent, n = -y

Ex: $0.00013482 = 1.3482\times10^{-4}$

$0.00549 = 5.49\times10^{-3}$

$0.1641 = 1.641\times10^{-1}$

➢ **SIGNIFICANT FIGURES**

- To express the results in an accurate way, we express generally those digits which are known with certainty. This is done in terms of significant figures.
- **The significant figures in a number are all the certain digits plus one doubtful digit**. The **digits in a properly recorded measurement are knows as significant figures**.
- The greater the number of significant figures in a reported result, smaller is the uncertainty and greater then precision.

➢ **Rules for determining number of significant figures:**

1. All non-zero digits are significant.

Ex: The number of significant figures in 1.887 = 4

Ex: The number of signficant figures in 12.612 = 5

Ex: The number of signficant figures in 1.23 = 3

2. When a number is greater than 1, all the zeros to the right of the decimal point are significant

Ex: The number of significant figures in 3.0 = 2

Ex: The number of signficant figures in 91.070 = 5

Ex: The number of significant figures in 42.000 = 5

3. For a number less than 1, only zeros to the right of the first significant digit are significant. But

the zeros to the left of the first significant digit are not significant

Ex: The number of significant figures in 0.4960 = 4
Ex: The number of significant figures in 0.0013 = 2
Ex: The number of significant figures in 0.0002 = 1
Ex: The number of significant figures in 0.030 = 2

4. **A zero becomes significant if it comes in between two non - zero digits**

   Ex: The number of significant figures in 3.01 = 3
   Ex: The number of significant figures in 6.023 = 4
   Ex: The number of significant figures in 3.0023 = 5

5. **When adding or subtracting**, the **number of decimal places in the answer should be equal to the number of decimal places in the number with the least number of decimal places**.

   Ex: 3.21(3 significant figures 2 decimal places)
   Ex: 1.5 (2 significant figures 1 decimal places)
   Ex: 21.402 (5 significant figures 3 decimal places)
   Since the term 1.5 involved in addition, has only one decimal place, the overall answer of 26.112 should be reported as 26.1.

6. **In multiplication and division, the number of significant figures in the answer should be same as that in the number with least number of significant figures**.

   Ex: Since the term 3.376 has 4 and 1.25 has 3 significant figures, the multiplied answer should be 4.22

7. **When a number is rounded off the number of significant figures is reduced.**

   Ex: If digit to be dropped is greater than 5, the lost retaining digit is increased by 1. Exp- 12.6 is rounded to 13. If digit to be dropped is less than 5, the lost retaining digit is left as it is. Exp- 12.4 is rounded to 12. If digit to be dropped is 5, the lost remaining digit is increased by 1, if it is odd but left as it is if even. Exp- 11.5 is rounded to 12 and 12.5 is rounded to 12last digit retained is increased by 1, only if the following digit is 5, and is left as such if the following digit is 4.

   Ex: 12.696, 18.35 and 13.93 are reported as 12.7, 18.4, 13.9 respectively when rounded off to three significant figures

MOLE CONCEPT

➢ In SI Units we represent mole by the symbol 'mol'. It is defined as follows:

(i) **A mole is the amount of a substance that contains as many entities (atoms, molecules or other particles) as there are atoms in exactly 12g of the carbon - 12 isotope.**

$$\frac{12\text{g / mol}\,C^{12}}{1.992648\times10^{-23}\,\text{g / }C^{12}\text{atom}} = 6.0221367\times10^{23}\ \text{atoms / mol}$$

1 MOLE
- One gram molecule of a compound
- One gram atom of a compound
- 6.023×10^{23} particles
- 22.4 litres of gas at NTP STP

(ii) **In a simple way, we can say that mole has 6.0221367×10^{23} entities (atom, molecules or ions etc.)** The number of entities in 1 mol is so important that it is given a separate name and symbol, known as **'Avogadro constant'** denoted by N_A •

Formula to get moles are following :

In number of species are given	If Mass is given	In volume is given
Number of moles $(n) = \frac{N}{N_A}$ (Where N = Number of particles) $N_A = 6.023 \times 10^{23}$	Number of moles $(n) = \frac{W(g)}{M.M.}$ (Where W = Mass of substance in (g) M.M.=Molar Mass	Number of moles $(n) = \frac{V_{(L)}}{22.4}$ (Where V = Volume of gas in L at NTP or STP)

❖ SOME RELATED DEFINITIONS :

(1) Atomic mass unit (a.m.u) or reference of mass: It is equal to $\frac{1}{12}$ mass of one atom of carbon –12 isotope.

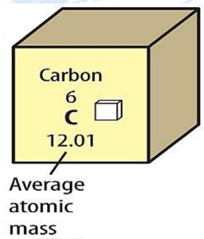
- 1 a.m.u. = $\frac{1}{12}$ × mass of one C–12 atom = $\frac{1.9924\times10^{-23}}{12}$ g = 1.66×10^{-24} g or 1.66×10^{-27} kg
- Today, a.m.u. is replaced by 'u' (Unified mass)

(2) relative atomic mass (R.A.M.)

$$= \frac{\text{mass of one atom of an element}}{\text{Reference mass}}$$

- R.A.M. is dimension less quantity which is the sum of number of protons and neutrons in the species.
  Ex. R.A.M. of hydrogen is 1 and R.A.M. of oxygen is 16.

(3) Average atomic mass: Elements are found in different isotopic forms (atoms of same elements having different atomic mass), so the atomic mass of any element is the average of all the isotopic mass within a given sample.



Average atomic mass =

$$\frac{\text{Mass of first isotope } (m_1) \times \text{Its\%}(x_1) + \text{Mass of 2nd isotope } (m_2) \times \text{Its\%}(x_2)}{\text{\% of first isotope} + \text{\% of 2nd isotope}}$$

or Average atomic mass = $\frac{m_1x_1 + m_2x_2}{100}$

Ex: Use the date given in the following table to calculate the molar mass of naturally occurring argon.

Isotope	Isotopic molar mass	Abundance
^{36}Ar	35.96755 g mol^{-1}	7.1%
^{38}Ar	37.96272g mol^{-1}	16.3%
^{40}Ar	39.9624 g mol^{-1}	76.6%

Sol: Molar mass of Ar
= 35.96755 × 0.071 + 37.96272 × 0.163 + 39.96924 × 0.766
= 39.352 g mol^{-1}

(4) Relative molecular mass (Molecular mass) : The number which indicates how many times the mass of one molecule of a substance is heavier in comparison to $\frac{1}{12}$th part of the mass of an atom of C–12

(5) Gram atomic mass: When R.A.M. of an element is expressed in gram or mass of one mole atoms of an element in gram.

gram atomic mass = mass of 1 **gram atom** = mass of 1 **mole atom** = mass of N_A atoms = mass of 6.023×10^{23} atoms.

Ex: Gram atomic mass of oxygen = mass of 1 **g atom** of oxygen = mass of 1 **mol atom** of oxygen.
= mass of N_A atoms of oxygen.
$= \left(\frac{16}{N_A} g\right) \times N_A = 16g$

(6) **Gram Molecular Mass (Mass of 1 Gram Molecule)/Molar Mass:** When numerical value of molecular mass of the substance is expressed in grams then the value becomes gram molecular mass.
gram molecular mass = mass of 1 **gram molecule**
= mass of 1 **mole molecule**
= mass of N_A molecules
= mass of 6.023×10^{23} molecules

Ex: Gram molecular mass of H_2SO_4 = mass of 1 **gram molecule** of H_2SO_4 =mass of 1 **mole molecule** of H_2SO_4
= mass of N_A molecules of H_2SO_4
$= \left(\frac{98}{N_A} g\right) \times N_A = 98g$

(7) **Actual Mass:**

- The mass of one atom or one molecule of substance expressed in gram is called as actual mass.

Ex:
(i) Actual mass of O_2 = 32 amu
= $32 \times 1.67 \times 10^{-24}$g → Actual mass
(ii) Actual mass of H_2O = (2 + 16) amu
= $18 \times 1.67 \times 10^{-24}$g = 2.99×10^{-23}g

(8) **GRAM MOLECULAR VOLUME (GMV)**

At NTP, the volume of 1 mole of gaseous substance is 22.4 litre is called as gram molecular volume.
At NTP, d_{H_2} = 0.000089 g/mL = mass/volume = mass/1000 mL
If volume = 1 litre= 1000 mL then mass = 0.089 g
∴ 0.089g H_2 occupies = 1 litre at STP.
∴ 1 g H_2 occupies = $\frac{1 \text{ litre}}{0.089}$ at STP.
∴ 2 g or 1 mol H_2 occupies = $\frac{1 \text{ litre}}{0.089} \times 2$ = 22.4 litre at STP

1 mole of any gaseous substance occupy 22.4 litre of volume at NTP or STP

$$\boxed{1 \text{ mol} = 22.4 \text{ litre } (\text{at STP})}$$

(9) Atomicity

- Total number of atoms in a **molecule** of elementary substance is called as atomicity.

Ex:

Molecule	Atomicity
H_2	2
O_2	2
O_3	3
NH_3	4

Ex: Calculate the number of molecules of sulphur dioxide in 0.064 g of the gas.

Sol: Gram molecular weight of sulphur dioxide (SO_2) = 64g

Given mass = 0.064 g

Gram molecular weight of any gas contain avogadro number of molecules = 6.023×10^{23}

$\therefore$ 0.064 g of sulphur dioxide contain $\left(\dfrac{6.023\times10^{23}}{1000}\right)$ molecules = 6.023×10^{20}

Ex: Which of the following contains the least number of molecules -

(1) 16g of CO_2 (2) 8g of O_2

(3) 4g of N_2 (4) 2g of H_2

Ans: (3)

Sol: (1) No. of moles of $CO_2 = \dfrac{\text{Weight}}{\text{Molecule weight}}$

$= \dfrac{16}{44} = 0.36$

(2) Number of moles of $O_2 = \dfrac{8}{32} = 0.25$

(3) Number of moles of $N_2 = \dfrac{4}{28} = 0.14$

(4) Number of moles of $H_2 = \dfrac{2}{2} = 1$

Ex: Atomic weight of helium is 4. Calculate the number of atoms in 1g of helium -

Sol: 4g of Helium contains 6.023×10^{23} atoms

1g of Helium contains $\dfrac{6.023\times10^{23}}{4}$

$= 1.506 \times 10^{23}$ atoms

Ex: What is the mass of 1 molecule of CO -

Sol: Gram molecular weight of CO = 12 + 16 = 28g

6.023×10^{23} molecules of CO weight 28g

1 molecule of CO weight $= \dfrac{28}{6.02\times10^{23}} = 4.65\times10^{-23}$ g

Ex: Calculate the volume at STP occupied by 240g of SO_2.

Sol: Molecular weight of $SO_2 = 32 + 2 \times 16 = 64$

64 g of SO_2 occupies 22.4 litre at STP

240 g of SO_2 occupies $= \dfrac{22.4}{64} \times 240 = 84$ litre at STP

Ex: Calculate the number of atoms in each of the following -

(1) 52 mole of He (2) 52 amu of He

(3) 52 g of He

Sol: (a) 1mole of He contain 6.02×10^{23} atoms

$\therefore$ 52 mole of He contain $= 52 \times 6.02 \times 10^{23}$

$= 31.3 \times 10^{24}$ atoms

(b) Atomic weight of He = 4amu

$\therefore$ 52 amu of He contain $= \dfrac{52}{4}$ = 13 atoms of He

(c) Number of moles of He in 52g $= \dfrac{52}{4} = 13$ moles

$\therefore$ no. of atoms in 52g of He i.e. 13 moles

$= 13 \times 6.02 \times 10^{23}$ atoms $= 78.26 \times 10^{23}$ atoms

(10) Degree Of Dissociation (α):

Degree of dissociation represents the fraction of one mole dissociated into the products.

(Defined for one mole of substance)

So, α = no. of moles dissociated / initial no. of moles taken

= fraction of moles dissociated out of 1 mole.

Note : % dissociation = $\alpha \times 100$

Suppose 5 moles of PCl_5 is taken and if 2 moles of PCl_5 dissociated then $\alpha = \dfrac{2}{5} = 0.4$

(11) Relationship Between Average Molar Mass & Degree Of Dissociation (α):

Let a gas A_n dissociates to give n moles of A as follows-

	An (g)	$\rightleftharpoons$	n A (g)
t = 0	a		0
$t = t_{eq}$	a – x		nx

$\alpha = \dfrac{x}{a} \Rightarrow x = a\alpha.$

$a - a\,\alpha = a(1 - \alpha)$ $\qquad$ $n\,a\,\alpha$

Total no. of moles $= a - a\,\alpha + n\,a\,\alpha = [1 + (n - 1)\,\alpha]\,a$

Average molecular weight of mixture(g)

$$= \frac{\text{molecular weight of } A_n(g)}{\text{total no. of moles at equilibrium}}$$

$$= \frac{a \cdot M_{th}}{a(1+(n-1)\alpha)}$$

$$M_{avg} = \frac{M_{th}}{[1+(n-1)\alpha]}$$

where M_{th} = theoritical molecular weight

(n = atomicity)

$$M_{mixture} = \frac{M_{A_n}}{[1+(n-1)\alpha]}, M_{A_n} = \text{Molar mass of gas } A_n$$

(12) DENSITY:

It is of two type.

1. Absolute density
2. Relative density

(a) For liquid and solids

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

Relative density or specific gravity

$$= \frac{\text{density of the substance}}{\text{density of water of 4°C}}$$

(b) For gases:

Absolute density (mass / volume) = $\frac{PM}{RT}$

where P is pressure of gas, M = mol. wt. of gas, R is the gas constant, T is the temperature.

Relative density and Vapour density:

Vapour density is defined as the density of the gas with respect to hydrogen gas at the same temperature and pressure.

$$\text{Vapour density} = \frac{d_{gas}}{d_{H_2}} = \frac{PM_{gas}/RT}{PM_{H_2}/RT}$$

$$\text{V.D.} = \frac{M_{gas}}{M_{H_2}} = \frac{M_{gas}}{2} \Rightarrow M_{gas} = 2\text{V.D.}$$

Vapour density (V.D). :

Density of the gas divided by density of hydrogen under same temperature & pressure is called vapour density.

D = vapour density without dissociation = $\frac{M_{A_n}}{2}$

d = vapour density of mixture = average vapour density

$$\frac{D}{d} = 1+(n-1)\alpha$$

$$\alpha = \frac{D-d}{(n-1)\times d} = \frac{M - M_o}{(n-1)M_o}$$

Ex: NH_3 decomposes into N_2 & H_2. If average molar mass of reaction mixture is 10 then, find α ?

Ans: 0.7

Sol.

NH_3	$\longrightarrow$	$\frac{1}{2}N_2$	+ $\frac{3}{2}H_2$
$n_1 1$		0	0
$1-\alpha$		$\frac{\alpha}{2}$	$\frac{3\alpha}{2}$

$$10 = \frac{17}{1-\alpha+\frac{\alpha}{2}+\frac{3\alpha}{2}} \qquad 10 = \frac{17}{1+\alpha} \qquad 1+\alpha = 1.7$$

$\alpha = 0.7$

Ex: Find the relative density of SO_3 gas with respect to methane:

(1) 8 (2) 3.5 (3) 2.5 (4) 5

Ans: (4)

Sol: R.D. $= \frac{M_{SO_3}}{M_{CH_4}} = \frac{80}{16} = 5$

Ex: The atomic mass of a metal is 27 u. If its valency is 3, the vapour density of the volatile metal chloride will be:

(1) 66.75 (2) 32.1

(3) 26.7 (4) 80.25

Ans: (1)

Sol: Element must be Al

Hence, volatile chloride will be $AlCl_3$ so V.D.

$$= \frac{M_{AlCl_3}}{2} = \frac{133.5}{2} = 66.75$$

Ex: The density of water at 4°C is 1×10^3 kg m^{-3}. Assuming no empty space to be present between water molecules, the volume occupied by one molecule of water is approximately:

(1) 3×10^{-23} mL (2) 6×10^{-23} mL

(3) 3×10^{-22} mL (4) 6×10^{-22} mL

Ans: (1)

Sol: 1×10^3 kg/m^3 = 1 g/mL.

[Since, 1m^3 = 10^6 cm^3 = 10^6 mL].

= 1 g/cc

6.022×10^{23} H_2O molecule weigh …18 g

1 H_2O molecule weigh… $\frac{18}{6.022\times10^{23}}\text{g} = 3\times10^{-23}\text{g}$

$d = \frac{\text{mass}}{\text{volume}}$,

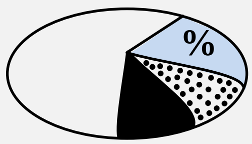
So, volume = $\frac{3\times10^{-23}\text{g}}{1(\text{g}/\text{mL})} = 3\times10^{-23}\text{mL}$

PERCENTAGE COMPOSITION, EMPIRICAL FORMULA & MOLECULAR FORMULA

PERCENTAGE FORMULA (% BY MASS) of any element or constituent in a compound is the number of parts by mass of that element or constituent present in 100 parts by mass of the compound.

Percent composition

The percentage of by mass of each element in a compound



Mass % of an element

$$= \frac{\text{Mass of an element in 1 mole of compound}}{\text{Mass of 1 mole compound}} \times 100$$

It can be calculated by the following two steps:

Step-1. Calculate the molecular mass of the compound from its formula by adding the atomic masses of the elements present.

Step-2. Calculate the percentage of the element or the constituent by apply the following relation:

% of the element

$$= \frac{\textit{No. of parts by mass of the elements or constituent}}{\textit{Mol.mass of the compound}} \times 100$$

Ex. : Calculate the percentage compositions of the various elements in $MgSO_4$.

Sol: Mol. mass of $MgSO_4$ = 24 + 32 + 4 x 16 = 120

$$\text{\% of Mg} = \frac{\text{No. of parts by mass of Mg}}{\text{Mol. Mass of MgSO}_4} \times 100$$

$$= \frac{24}{120} \times 100 = 20\%$$

$$\text{\% of S} = \frac{\text{No. of parts by mass of S}}{\text{Mol. mass of MgSO}_4} \times 100$$

$$= \frac{32}{120} \times 100 = 26.67\ \%$$

$$\text{\% of O} = \frac{\text{No. of parts by mass of O}}{\text{Mol. mass of MgSO}_4} \times 100$$

$$= \frac{64}{120} \times 100 = 53.33\ \%$$

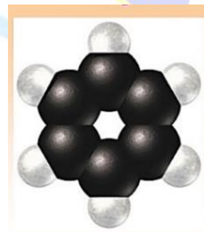
Empirical formula

An empirical formula is the simplest whole number ratio of elements in a compound.

MolecularFormula	**Empirical formula**
C_2H_6	CH_3
C_4H_8	CH_2
$C_6H_{12}O_6$	CH_2O

Molecular Formula

The molecular formula of a compound represents the actual number of atoms present in 1 molecule of the compound i.e. it shows the real formula of its 1 molecule.



BENZENE

Molecular formula : C_6H_6

Empirical Formula : CH

Relationship between Empirical & Molecular Formula

Molecular Formula = n × Empirical Formula

[Where n = natural no. (1, 2, 3, ...)]

$$\text{Or } n = \frac{\text{Molecular Formula}}{\text{Empirical Formula}}$$

$$\text{Or } n = \frac{\text{Molecular Formula mass}}{\text{Empirical Formula mass}}$$

➢ **Determination of Empirical Formula**

Following steps are involved to determine the empirical formula of the compounds -

Step-1 First of all find the % by weight of each element present in 1 molecule of the compound.

Step-2 The % by weight of each element is divided by its atomic weight. It gives atomic ratio of elements present in the compounds.

Step-3 Atomic ratio of each element is divided by the minimum value of atomic ratio as to get simplest ratio of atoms.

Step-4 If the value of simplest atomic ratio is fractional then raise the value to the nearest whole number. or Multiply with suitable coefficient to convert it into nearest whole number.

Step-5 Write the Empirical formula as we get the simplest ratio of atoms.

Ex: Phosgene, a poisonous gas used during World war-I, contains 12.1% C, 16.2% O and 71.7% Cl by mass. What is the empirical formula of phosgene.

Sol.

Element	**%**	**Mole ratio**	**Simplest mole ratio**
C	12.1	$\frac{12.1}{12} = 1.01$	$\frac{1.01}{1.01} = 1$
O	16.2	$\frac{16.2}{16} = 1.01$	$\frac{1.01}{1.01} = 1$
Cl	71.7	$\frac{71.7}{35.5} = 2.02$	$\frac{2.02}{1.01} = 2$

Then empirical formula = $COCl_2$

Ex: 5.325g sample of methyl benzoate, a compound used in the manufacture of perfumes is found to contain 3.758 g of carbon, 0.316g hydrogen and 1.251g of oxygen. What is empirical formulae, of compound. If mol. weight of methyl benzoate is 136.0, calculate its molecular formula.

Sol:

Element	**%**	**Mole ratio**	**Simplest whole ratio**
C	$\frac{3.758 \times 100}{5.325} = 70.57$	$\frac{70.57}{12} = 5.88$	$\frac{5.88}{1.47} = 4$
H	$\frac{0.316 \times 100}{5.325} = 5.93$	$\frac{5.93}{1} = 5.93$	$\frac{5.93}{1.47} = 4$
O	$\frac{1.251 \times 100}{5.325} = 23.50$	$\frac{23.50}{16} = 1.47$	$\frac{1.47}{1.47} = 1$

Empirical = C_4H_4O

$$n = \frac{\text{Mol. wt}}{\text{Empirical formula wt.}} = \frac{136}{68} = 2$$

$\Rightarrow$ Molecular formula = $C_8H_8O_2$

SPOT LIGHT

BALANCING A CHEMICAL EQUATION

According to the law of conservation of mass, a balanced chemical equation has the same number of atoms of each element on both sides of the equation. Many chemical equations can be balanced by trial and error. Let us take the reactions of a few metals and non-metals with

Ex. Combustion of propane, C_3H_8. This equation can be balanced in steps.

Step 1 Write down the correct formulas of reactants and products. Here, propane and oxygen are reactants, and carbon dioxide and water are products.

$C_3H_8(g) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$ unbalanced equation

Step 2 Balance the number of C atoms: Since 3 carbon atoms are in the reactant, therefore, three CO_2 molecules are required on the right side.

$C_3H_8(g) + O_2(g) \rightarrow 3CO_2(g) + H_2O(l)$

Step 3 Balance the number of H atoms: on the left there are 8 hydrogen atoms in the reactants however, each molecule of water has two hydrogen atoms, so four molecules of water will be required for eight hydrogen atoms on the right side.

$C_3H_8(g) + O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$

Step 4 Balance the number of O atoms: There are 10 oxygen atoms on the right side ($3 \times 2 = 6$ in CO_2 and $4 \times 1 = 4$ in water). Therefore, five O_2 molecules are needed to supply the required 10 CO_2 and $4 \times 1 = 4$ in water).

$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$

Step 5 Verify that the number of atoms of each element is balanced in the final equation. The equation shows three carbon atoms, eight hydrogen atoms, and 10 oxygen atoms on each side.

All equations that have correct formulas for all reactants and products can be balanced. Always remember that subscripts in formulas of reactants and products cannot be changed to balance an equation.

CHEMICAL EQUATION

Representation of the chemical change in terms of symbol and formulae of the reactants & products is called a chemical equation.

➢ **Information conveyed by a chemical equation**

(1) Qualitatively, a chemical equation tells us the names of the various reactants

(2) Quantitatively, it express
- (a) physical states of reactant & products.
- (b) The relative no. of molecules of reactants and products
- (c) The relative no. of moles of reactant and products
- (d) The relative masses of reactants and products
- (e) The relative volumes of gaseous reactants and products

STOICHIOMETRY AND STOICHIOMETRIC CALCULATIONS

The word 'stoichiometry' is derived from two Greek words — *stoicheion*(meaning, *element*) and *metron* (meaning, *measure*).Stoichiometry, thus, deals with the calculation of masses (sometimes volumes also) of the reactants and the products involved in a chemical reaction

Ex: When potassium chlorate ($KClO_3$) is heated it gives potassium chloride (KCl) and oxygen (O_2).

$$KClO_3 \xrightarrow{\Delta} KCl + O_2$$

(unbalanced chemical equation)

$$2KClO_3 \xrightarrow{\Delta} 2KCl + 3O_2$$

(balanced chemical equation)

- Remember a balanced chemical equation is one which contains an equal number of atoms of each element on both sides of equation.

1. Interpretation of balanced chemical equations:

Once we get a balanced chemical equation then we can interpret a chemical equation by following ways

(a) Mole - mole analysis
(b) Mass - mass analysis
(c) Mass - volume analysis

(a) Mole - mole analysis:

This analysis is very much important for quantitative analysis point of view.
Now consider again the decomposition of $KClO_3$.

$$2\,KClO_3 \longrightarrow 2\,KCl + 3O_2$$

In very first step of mole-mole analysis you should read the balanced chemical equation like **2 moles $KClO_3$ on decomposition gives you 2 moles KCl and 3 moles O_2** and from the stoichiometry of reaction we can write

$$\frac{\text{Moles of } KClO_3}{2} = \frac{\text{Moles of KCl}}{2} = \frac{\text{Moles of } O_2}{3}$$

Now for any general balance chemical equation like

$$a\,A + b\,B \longrightarrow c\,C + d\,D$$

you can write.

$$\frac{\text{Moles of A reacted}}{a} = \frac{\text{Moles of B reacted}}{b}$$

$$= \frac{\text{Moles of C reacted}}{c} = \frac{\text{Moles of D reacted}}{d}$$

(b) Mass - mass analysis :

Consider the reaction $2\,KClO_3 \longrightarrow 2KCl + 3O_2$

According to stoichiometry of the reaction

or $$\frac{\text{Mass of } KClO_3}{\text{Mass of KCl}} = \frac{2\times 122.5}{2\times 74.5}$$

$$\frac{\text{Mass of } KClO_3}{\text{Mass of } O_2} = \frac{2\times 122.5}{3\times 32}$$

Ex: Consider the balanced reaction

$2Cl_2O_7 \longrightarrow 4ClO_2 + 3O_2$ (Cl = 35.5)

What can be concluded from the coefficients of species in this balanced equation?

(1) For this reaction, exactly 2 g of Cl_2O_7 must be taken to start the reaction

(2) For this reaction, exactly 2 mol of Cl_2O_7 must be taken to start the reaction

(3) Mole ratio of Cl_2O_7, ClO_2 and O_2 during a chemical reaction at any instant are 2, 4 and 3 respectively

(4) The ratio of change in number of moles of Cl_2O_7, ClO_2 and O_2 is 2: 4: 3

Ans: (4)

Sol: It follows directly from definition of stoichiometry.

Ex: Calculate the weight of iron which will be converted into its oxide by the action of 36 g of steam.

(Given: $3Fe + 4H_2O \longrightarrow Fe_3O_4 + 4H_2$)

Ans: 84 g

Sol: Mole ratio of reaction suggests,

Mole of Fe = $\frac{3}{4}$

mol of H_2O = $\frac{3}{4}\times\frac{36}{18} = \frac{3}{2}$

wt. of Fe = $\frac{3}{2} \times 56 = 84$g

Ex: When Dinitrogen pentoxide (N_2O_5, a white solid) is heated, it decomposes into nitrogen dioxide and oxygen.

If a sample of N_2O_5 produces 1.6 g O_2, then how many grams of NO_2 are formed?

$N_2O_5(s) \longrightarrow NO_2(g) + O_2(g)$ (not balanced)

(1) 9.2 g (2) 4.6 g

(3) 2.3 g (4) 18.4 g

Ans: (1)

Sol: $N_2O_5(s) \longrightarrow 2NO_2(s) + \frac{1}{2}O_2$ (Balanced reaction)

$$\frac{\text{Mole of } O_2}{1/2} = \frac{\text{Mole of } NO_2}{2}$$

$$\frac{1.6}{32} \times 2 \times 2 = \text{Mole of } NO_2 = 0.2$$

wt. of $NO_2 = 0.2 \times 46 = 9.2$ g.

(c) Mass - volume analysis:

Now again consider decomposition of $KClO_3$

$2\ KClO_3 \longrightarrow 2KCl + 3O_2$

mass volume ratio: 2 × 122.5 g: 2 × 74.5 g: 3 × 22.4 L at STP

we can use two relation for volume of oxygen

$$\frac{\text{Mass of } KClO_3}{\text{Volume of } O_2 \text{ at STP}} = \frac{2 \times 122.5\text{g}}{3 \times 22.4\text{L}} \quad \ldots\ldots\text{(i)}$$

And

$$\frac{\text{Mass of KCl}}{\text{Volume of } O_2 \text{ at STP}} = \frac{2 \times 74.5\text{g}}{3 \times 22.4\text{L}} \quad \ldots\ldots\text{(ii)}$$

Ex: When oxygen gas is passed through Siemen's ozoniser, it completely gets converted into ozone gas. The volume of ozone gas produced at 1 atm and 273K, if initially 96 g of oxygen gas was taken, is:

(1) 44.8 L (2) 89.6 L

(3) 67.2 L (4) 22.4 L

Ans: (1)

Sol: $3O_2 \longrightarrow 2O_3$

Mole = $\frac{96}{32} = 3$ mole = 2

Volume of O_3 gas at 1 atm and 273K

= 2 × 22.4 = 44.8 L

LIMITING REAGENT

- Definition: It may be defined as the reactant which is completely consumed during the reaction is called limiting reagent-
- Many a time, reactions are carried out with the amounts of reactants that are different than the amounts as required by a balanced chemical reaction. In such situations, one reactant is in more amount than the amount required by balanced chemical reaction.
- The reactant which is present in the least amount gets consumed after sometime and after that further reaction does not take place whatever be the amount of the other reactant. Hence, the reactant, which gets consumed first, limits the amount of product formed and is, therefore, called the limiting reagent.
- The reactant which consumed first into the reaction when we are dealing with balance chemical equation then if number of moles of reactants are not in the ratio of stoichiometric coefficient of balanced chemical equation, then there should be one reactant which should be limiting reactant.

Ex: $2H_2(g) + O_2(g) \rightarrow 2H_2O\ (l)$

Reaction is started with one mole each of H_2 and O_2

Here H_2 is known as limiting reagent.

Ex: Six mole of Na_2CO_3 is reacted with 4 moles of HCl solution. Find the volume of CO_2 gas produced at STP. The reaction is

$Na_2CO_3 + 2\ HCl \longrightarrow 2\ NaCl + CO_2 + H_2O$

Sol: From the reaction:

$Na_2CO_3 + 2\ HCl \longrightarrow 2\ NaCl + CO_2 + H_2O$

Given moles 6 mol 4 mol

given mole ratio 3 : 2

Stoichiometric coefficient ratio 1 : 2

- See here given number of moles of reactants are not in stoichiometric coefficient ratio.
- Therefore there should be one reactant which consumed first and becomes limiting reagent.
- But the question is how to find which reactant is limiting, it is not very difficult you can

easily find it. According to the following method.

HOW TO FIND LIMITING REAGENT:

Step: I Divide the given moles of reactant by the respective stoichiometric coefficient of that reactant.

Step: II See for which reactant this division comes out to be minimum. The reactant having minimum value is limiting reagent.

Step: III Now once you find limiting reagent then your focus should be on limiting reagent

From Step I & II

Na_2CO_3 $\quad$ HCl

$\frac{6}{1} = 6$ $\quad$ $\frac{4}{2} = 2$ (Division in minimum)

$\therefore$ HCl is limiting reagent

From Step III

From

$$\frac{\text{Mole of HCl}}{2} = \frac{\text{Mole of } CO_2 \text{ produced}}{1}$$

$\therefore$ Mole of CO_2 produced = 2 moles

$\therefore$ Volume of CO_2 produced at S.T.P. = 2 × 22.7 = 45.4 L

Ex: In the reaction $4A + 2B + 3C \longrightarrow A_4B_2C_3$ what will be the number of moles of product formed, starting from 2 moles of A, 1.2 moles of B & 1.44 moles of C :

(1) 0.5 $\quad$ (2) 0.6

(3) 0.48 $\quad$ (4) 4.64

Ans: (C)

	4A +	2B +	3C ⟶	$A_4B_2C_3$
Initial mole	2	1.2	1.44	0
final mole			0	0.48

C is limiting reagent.

moles of $A_4B_2C_3$ is 0.48.

EXAMPLES BASED ON CHEMICAL REACTIONS

Ex: Calculate the mass of oxygen required to burn 14g C_2H_4 completely-

Sol:

	C_2H_4 +	$3O_2$ →	$2CO_2$ +	$2H_2O$
Mole ratio	1	3	2	4

Moles of C_2H_4 to be burnt $= \frac{14}{28} = \frac{1}{2}$ mole.

$\because$ 1 mole C_2H_4 requires 3 mol O_2 for combustion

$\therefore$ $\frac{1}{2}$ C_2H_4 requires $3 \times \frac{1}{2}$ mole $O_2 = \frac{3}{2}$ mol O_2.

& Mass of $O_2 = \frac{3}{2} \times 32 = 48$ g

Ex: Calculate the weight and volume of H_2 at STP that will be displaced by 1 gram of Zn when it is completely dissolved in dilute sulphuric acid.

Sol: $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$

1 atom	1 molecule
1 molecule	1 molecule
1 gram-atom	65.4 g
1 mole	2 g or 22.4 dm^3

$\because$ 65.4 g of Zn displaces 2g of Hydrogen

$\therefore$ 1.0 g of Zn displaces $\frac{2}{65.4} \times 1 = 0.0306$ g of H_2

$\because$ 65.4 g of Zn displaces 22.4 dm^3 of H_2 at S.T.P.

$\therefore$ 1.0 g of Zn displaces $\frac{22.4}{65.4} \times 1.0 = 0.3425\ dm^3$

Ex: 10 ml of liquid carbon disulphide (sp. gravity 2.63) is burnt in oxygen. Find the volume of the resulting gases measured at STP.

Sol: 1 ml of CS_2 Weighs 2.63 g

10 ml of CS_2 weighs 26.3 g

$CS_2 + 3O_2 \rightarrow$	$CO_2 + 2SO_2$
12 + (2 × 32)	22.4 ℓ $\quad$ 44.8 ℓ
76 g	67.2ℓ.

$\because$ 76g of CS_2 will yield 67.2 L of a mixture of CO_2 and SO_2 at STP

$\therefore$ 26.3 g of CS_2 would yield $\frac{67.2}{76}$a × 26.3

= 23.26 lit.

EQUIVALENT WEIGHT

The equivalent weight of a substance is the number of parts by weight of the substance that combine displace directly or indirectly 1.008 parts by weight of hydrogen or 8 parts by weight of oxygen or 35.5 parts by weight of chlorine or 108 parts by weight of Ag.

(a) Calculation Of Equivalent Weight

(i) Equivalent weight $= \frac{\text{Atomic weight}}{\text{Valency factor}}$

(ii) Equivalent weight of ions

$= \frac{\text{formula weight of ion}}{\text{Charge on ion}}$

(iii) Equivalent weight of ionic compound = equivalent weight of cation + equivalent weight of anion

(iv) Equivalent weight of acid / base

$$= \frac{\text{Molecular weight}}{\text{Basicity / Acidicity}}$$

(v) Equivalent weight of salt

$$= \frac{\text{Molecular weight}}{\text{Total charge on cation or anion}}$$

(vi) Equivalent weight of an oxidizing or reducing agent

$$= \frac{\text{Molecular weight of the substance}}{\text{Number of electrons gain/lost by one molecule}}$$

Ex. Equivalent weight of H_2SO_4 = Equivalent weight of H^+ + Equivalent weight of Anion $\left(SO_4^{-2}\right)$

= 1 + 48 = 49

Ex. $Na_2SO_4(\text{salt}) \rightarrow 2Na^+ + SO_4^{-2}$

Total charge on cation or anion is 2

molecular weight of Na_2SO_4 is = (2 × 23 + 32 + 16 × 4) = 142

Equivalent weight of $Na_2SO_4 = \frac{142}{2} = 71$

(b) Concept Of Gram Equivalent And Law Of Chemical Equivalence

Number of gram equivalent

$$= \frac{W_{(g)}}{E} \Rightarrow \frac{W_{(g)}}{\frac{\text{Atomic weight}}{n-\text{factor}}} = \frac{W_{(g)}}{\text{Atomic weight}} \times n-\text{factor}$$

= mole × valence factor ; where

$$\left(\text{Normality} = \frac{\text{number of gram equivalent of solute}}{\text{volume of solution in } (L)}\right)$$

According to it in a reaction equal gram equivalent of reactant reacts to give same number of gram equivalent of products.

$aA + bB \rightarrow cC + dD$

Number of gram equivalent of A = Number of gram equivalent of B = Number of gram equivalent of C = Number of gram equivalent of D

(c) Methods For Determination Of The Equivalent Weights

(i) Hydrogen displacement method : This method is used for those elements which can evolve hydrogen from acids, i.e., active metals.

equivalent weight of metal

$$= \frac{\text{weight of metal}}{\text{weight of } H_2 \text{ gas (displaced)}} \times 1.008$$

(ii) Oxide formation method : A known mass of the element is changed into oxide directly or indirectly. The mass of oxide is noted weight of oxygen = weight of oxide – weight of element

equivalent weight of element

$$= \frac{\text{Weight of element}}{\text{weight of oxygen}} \times 8$$

(iii) Chloride formation method : A known mass of the element is changed into chloride directly or indirectly. The mass of the chloride is determined.

equivalent weight of element

$$= \frac{\text{weight of element}}{\text{weight of chlorine}} \times 35.5$$

(iv) Metal to metal displacement method : More active metal can displace less active metal from its salt's solution. The mass of the displaced metal bear the same ratio as their equivalent weights.

$$\frac{m_1}{m_2} = \frac{E_1}{E_2}$$

(v) Double decomposition method: this method is based on the following points -

(a) The mass of the compound reacted and the mass of product formed are in the ratio of their equivalent masses.

(b) The equivalent mass of the compound (electrovalent) is the sum of equivalent masses of its radicals.

(c) The equivalent mass of a radical is equal to the formula mass of the radical *divided* by its charge.

$$= \frac{\text{Mass of AB}}{\text{Mass of AD}} = \frac{\text{Equivalent mass of AB}}{\text{Equivalent mass of AD}}$$

$$= \frac{\text{Equivalent mass of A+ Equivalent mass of B}}{\text{Equivalent mass of A + Equivalent mass of D}}$$

(vi) Silver salt method : This method is used for finding the equivalent weight of carboxylic (organic) acids. A known mass of the RCOOAg is changed into Ag through combustion. The mass of Ag is determined.

$$\frac{\text{Equivalent weight of RCOOAg}}{\text{Equivalent weight of Ag}} = \frac{\text{weight of RCOOAg}}{\text{weight of Ag}}$$

Equivalent weight of RCOOAg

$$= \frac{\text{weight of RCOOAg}}{\text{weight of Ag}} \times 108$$

(vii) By electrolysis : $\frac{w_1}{w_2} = \frac{E_1}{E_2}$

Where w_1& w_2 are deposited weight of metals at electrodes and E_1 and E_2 are equivalent weight respectively.

CONCENTRATION TERMS

Solutions :

A mixture of two or more substances can be a solution. We can also say that "a solution is a homogeneous mixture of two or more substances,'' 'Homogeneous' means 'uniform throughout'. Thus a homogeneous mixture, i.e., a solution, will have uniform composition throughout the solution

➢ **Concentration terms :**

The following concentration terms are used to expressed the concentration of a solution. These are
(i) Molarity (M) (ii) Molality (m)
(iii) Mole fraction (x) (iv) % calculation
(v) ppm

❍ Remember that all of these concentration terms are related to one another. By knowing one concentration term you can also find the other concentration terms. Let us discuss all of them one by one.

Molarity (M) :

The number of moles of a solute dissolved in 1 L (1000 ml) of the solution is known as the molarity of the solution.

i.e., Molarity of solution

$$= \frac{\text{number of moles of solute}}{\text{volume of solution in litre}}$$

Let a solution is prepared by dissolving w g of solute of mol. wt. M in V ml water.

$\therefore$ Number of moles of solute dissolved $= \frac{w}{M}$

$\therefore$ V ml water have $\frac{w}{M}$ mole of solute

$\therefore$ 1000 ml water have $\frac{w \times 1000}{M \times V_{ml}}$

$\therefore$ Molarity (M) $= \frac{w \times 1000}{(\text{Mol. wt of solute}) \times V_{ml}}$

Some other relations may also useful.
Number of milli moles

$$= \frac{\text{mass of solute}}{(\text{Mol. wt. of solute})} \times 1000$$

$= (\text{Molarity of solution} \times V_{ml})$

❍ Molarity of solution may also given as :

$$M = \frac{\text{Number of millimole of solute}}{\text{Total volume of solution in ml}}$$

❍ Molarity is a unit that depends upon temperature. It varies inversely with temperature.

Mathematically : Molarity decreases as temperature increases.

$$\text{Molarity} \propto \frac{1}{\text{temperature}} \propto \frac{1}{\text{volume}}$$

❍ If a particular solution having volume V_1 and molarity M_1 is diluted upto volume V_2 mL than
$M_1V_1 = M_2V_2$
M_2 : Resultant molarity

❍ If a solution having volume V_1 and molarity M_1 is mixed with another solution of same solute having
volume V_2 mL & molarity M_2 then $M_1V_1 + M_2V_2 = M_R(V_1 + V_2)$

$$M_R = \text{Resultant molarity} = \frac{M_1V_1 + M_2V_2}{V_1 + V_2}$$

Ex: 149 g of potassium chloride (KCl) is dissolved in 10 L of an aqueous solution. Determine the molarity of the solution (K = 39, Cl = 35.5)

Sol: Molecular mass of KCl = 39 + 35.5 = 74.5 g

$\therefore$ Moles of KCl $= \frac{149\text{ g}}{74.5\text{ g}} = 2$

$\therefore$ Molarity of the solution $= \frac{2}{10} = 0.2$ M

Molality (m) :

The number of moles of solute dissolved in 1000 g (1 kg) of a solvent is known as the molality of the solution.

i.e., molality

$$= \frac{\text{number of moles of solute}}{\text{mass of solvent in gram}} \times 1000$$

Let Y g of a solute is dissolved in X g of a solvent. The molecular mass of the solute is M_0. Then Y/M_0 mole of the solute are dissolved in X g of the solvent. Hence Molality $= \frac{Y}{M_0 \times X} \times 1000$

❍ **Molality is independent of temperature changes.**

Ex: 255 g of an aqueous solution contains 5 g of urea. What is the concentration of the solution in terms of molality. (Mol. wt. of urea = 60)

Sol: Mass of urea = 5 g
Molecular mass of urea = 60

Number of moles of urea $= \frac{5}{60} = 0.083$

Mass of solvent = (255 – 5) = 250 g

$\therefore$ Molality of the solution

$$= \frac{\text{Number of moles of solute}}{\text{Mass of solvent in gram}} \times 1000$$

$= \frac{0.083}{250} \times 1000 = 0.332.$

Mole fraction (x) :

The ratio of number of moles of the solute or solvent present in the solution and the total number of moles present in the solution is known as the mole fraction of substances concerned.

Let number of moles of solute in solution = n

Number of moles of solvent in solution = N

$\therefore$ Mole fraction of solute $(x_1) = \frac{n}{n+N}$

$\therefore$ Mole fraction of solvent $(x_2) = \frac{N}{n+N}$

also $x_1 + x_2 = 1$

- Mole fraction is a pure number. It will remain independent of temperature changes.

% calculation :

The concentration of a solution may also expressed in terms of percentage in the following way.

- **% weight by weight (w/w) :** It is given as mass of solute present in per 100 g of solution.

  i.e. $\% \text{ w/w} = \frac{\text{mass of solute in g}}{\text{mass of solution in g}} \times 100$

- **% weight by volume (w/v) :** It is given as mass of solute present in per 100 ml of solution.

  i.e., $\% \text{ w/v} = \frac{\text{mass of solute in g}}{\text{volume of solution in ml}} \times 100$

- **% volume by volume (v/v) :** It is given as volume of solute present in per 100 ml solution.

  i.e., $\% \text{ v/v} = \frac{\text{volume of solute in ml}}{\text{volume of solution in ml}} \times 100$

Ex: 0.5 g of a substance is dissolved in 25 g of a solvent. Calculate the percentage amount of the substance in the solution.

Sol: Mass of substance = 0.5 g

Mass of solvent = 25 g

$\therefore$ Percentage of the substance (w/w)

$= \frac{0.5}{0.5+25} \times 100 = 1.96$

Ex: 20 cm^3 of an alcohol is dissolved in 80 cm^3 of water. Calculate the percentage of alcohol in solution.

Sol: Volume of alcohol = 20 cm^3

Volume of water = 80 cm^3

$\therefore$ Percentage of alcohol $= \frac{20}{20+80} \times 100 = 20.$

❖ **Parts Per Million (ppm)**

When the solute is present in very less amount, then this concentration term is used. It is defined as the number of parts of the solute present in every 1 million parts of the solution.

ppm can both be in terms of mass or in terms of moles. If nothing has been specified, we take ppm to be in terms of mass. Hence, a 100 ppm solution means that 100 g of solute is present in every 10,00,000 g of solution.

$\text{ppm} = \frac{\text{mass of A}}{\text{Total mass}} \times 10^6 = \text{mass fraction} \times 10^6$

❖ **Normality (N) :-**

The number of equivalent of a solute dissolved in 1 L (1000 ml) of the solution is known as the Normality of the solution.

i.e., Normality of solution

$= \frac{\text{number of equivalent of solute}}{\text{volume of solution in litre}}$

$\therefore$ Normality (N)

$= \frac{w \times 1000}{(\text{equivalent wt of solute}) \times V_{ml}}$

Normality (N) $= \frac{w \times 1000}{E \times V_{ml}}$

Ex: 8 litre of H_2 and 6 litre of Cl_2 are allowed to react to maximum possible extent. Find out the final volume of reaction mixture. Suppose P and T remains constant throughout the course of reaction –

(1) 7 litre (2) 14 litre

(3) 2 litre (4) None of these.

Sol: **(2)**

	H_2 +	Cl_2 $\rightarrow$	2 HCl
Volume before reaction	8 lit	6 lit	0
Volume after reaction	2	0	12

$\therefore$ Volume after reaction

= Volume of H_2 left + Volume of HCl formed

= 2 + 12 = 14 lit

Ex: Naturally occurring chlorine is 75.53% Cl^{35} which has an atomic mass of 34.969 amu and 24.47% Cl^{37} which has a mass of 36.966 amu. Calculate the average atomic mass of chlorine-

(1) 35.5 amu (2) 36.5 amu

(3) 71 amu (4) 72 amu

Sol: **(1)**

Average atomic mass

$= \frac{\% \text{ of I isotope} \times \text{its atomic mass} + \% \text{ of II isotope} \times \text{its atomic mass}}{100}$

$$= \frac{75.53 \times 34.969 + 24.47 \times 36.96}{100} = 35.5 \text{ amu.}$$

Ex: Calculate the mass in g of 2g atom of Mg-
(1) 12 g (2) 24 g
(3) 6 g (4) None of these.

Sol: (4)
$\because$ 1 g atom of Mg has mass = 24 g
$\therefore$ 2 g atom of Mg has mass = 24 x 2 = 48 g.

Ex: In 5 g atom of Ag (At. wt. of Ag = 108), calculate the weight of one atom of Ag -
(1) 17.93×10^{-23}g (2) 16.93×10^{-23}g
(3) 17.93×10^{23}g (4) 36×10^{-23}g

Sol: (1)
$\because$ N atoms of Ag weigh 108 g
$\therefore$ 1 atom of Ag weigh = $\frac{108}{N} = \frac{108}{6.023 \times 10^{23}}$
= 17.93×10^{-23} g.

Ex: In 5g atom of Ag (at. wt. = 108), calculate the no. of atoms of Ag –
(1) 1 N (2) 3N
(3) 5 N (4) 7 N.

Sol: (3)
$\because$ 1g atom of Ag has atoms = N
$\therefore$ 5g atom of Ag has atoms = 5N.

Ex: Calculate the mass in g of 2N molecules of CO_2-
(1) 22 g (2) 44 g
(3) 88 g (4) None of these.

Sol: (3)
$\because$ N molecules of CO_2 has molecular mass = 44.
$\therefore$ 2N molecules of CO_2 has molecular mass
= 44 x 2 = 88 g.

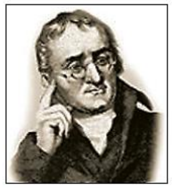
Ex: How many carbon atoms are present in 0.35 mol of $C_6H_{12}O_6$-
(1) 6.023×10^{23} carbon atoms
(2) 1.26×10^{23} carbon atoms
(3) 1.26×10^{24} carbon atoms
(4) 6.023×10^{24} carbon atoms

Sol: (3)
$\therefore$ 1 mol of $C_6H_{12}O_6$ has = 6 N atoms of C
$\therefore$ 0.35 mol of $C_6H_{12}O_6$ has
= 6 × 0.35 N atoms of C = 2.1 N atoms
= $2.1 \times 6.023 \times 10^{23} = 1.26 \times 10^{24}$ carbon atoms

Ex: How many molecules are in 5.23 g of glucose ($C_6H_{12}O_6$) -
(1) 1.65×10^{22} (2) 1.75×10^{22}
(3) 1.75×10^{21} (4) None of these

Sol: (2)
$\because$ 180 g glucose has = N molecules
$\therefore$ 5.23 g glucose has = $\frac{5.23 \times 6.023 \times 10^{23}}{180}$
= 1.75×10^{22} molecules

DALTON'S ATOMIC THEORY:



John Dalton (1776–1884)

- Dalton proposed the atomic theory on the basis of the law of conservation of mass and law of definite proportions. He also proposed the law of multiple proportion as a logical consequence of this theory. The salient features of this theory are

(a) Each element is composed by extremely small particles called atoms.

(b) Atoms of a particular element are all alike but differ with the atoms of other elements.

(c) Atom of each element is an ultimate particle, and has a characteristic mass but is structureless.

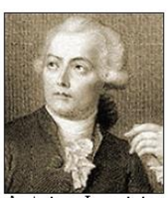
(d) Atom is indestructible i.e. it can neither be destroyed nor be created by any chemical reactions.

(e) Atom of an element takes part in chemical reaction to form molecule.

(f) Atoms of different elements combine in fixed ratio of small whole numbers to form compound (now called molecules).

LAWS OF CHEMICAL COMBINATION:

(a) Law of Mass Conservation (Law of Indestructibility of Matter)



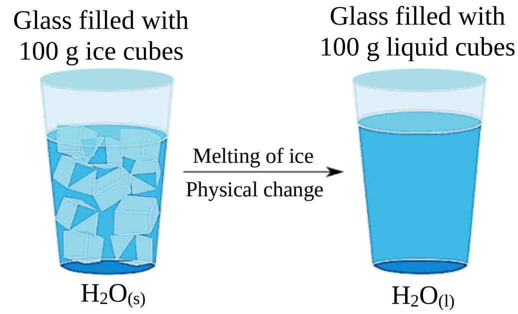
Antoine Lavoisier (1743–1794)

It states that "**matter is neither created nor destroyed during any physical or chemical change**"
This law is also called the Law of indestructibility of matter.

Thus,
Total mass of reactants = Total mass of products
(Before reaction) (After the chemical reaction)
Chemical combination is a must for the validity of this law.

- **When matter undergoes a physical change:** A piece of ice (solid water) is taken in a small conical flask. It is well corked and weighed. The flask is now heated gently to melt the ice (solid) into water (liquid).

$$\text{Ice} \xrightarrow{\text{Heat}} \text{Water}$$



The flask is again weighed. It is found that there is no change in the Weight though a physical change has taken place.

- **When matter undergoes a chemical change:** The following chemical changes illustrate the law.

Ex: Decomposition of Mercuric oxide: 100g of mercuric oxide when heated in a closed tube, decomposed to produce 92.6g of mercury and 7.4g of oxygen gas,
i.e. total mass of products = 100 g:

$$HgO(s) \rightarrow Hg(l) + \frac{1}{2}O_2(g)$$

100g 92.6g 7.4 g

Thus, in above decomposition reaction, matter is neither gained nor lost.

(b) Law of Definite Proportion / Law of Constant Composition

Joseph Proust (1754–1826)

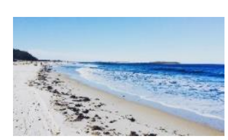
states that "**Any pure compound however made contains the same elements in the fixed ratio of their weights", or 'A pure chemical compound always contains the same elements combined together in the fixed ratio of their weights whatever its methods of preparation may be'.**

Ex: Pure water contains 2g of hydrogen and 16g of oxygen i.e., the ratio of hydrogen and oxygen in pure water is 1: 8.

Rain Water

River Water

Tap Water



Sea Water

Or we can say Water can be obtained from different sources but the ratio of weight of H and O remains same.

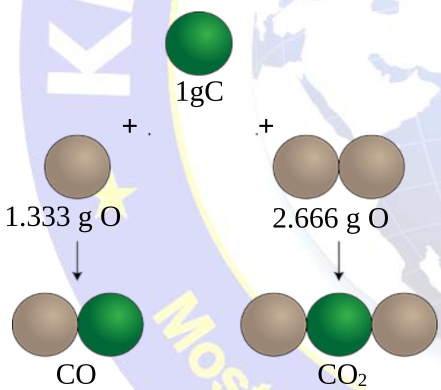
H_2		O		
2	:	16	←	← $2H_2O \leftarrow 2H_2 + O_2$
or 2	:	8		← Sea water
				← Tap water
				← River water
				← Ganga Jal
				← Rain Water

(c) Law of Multiple Proportion

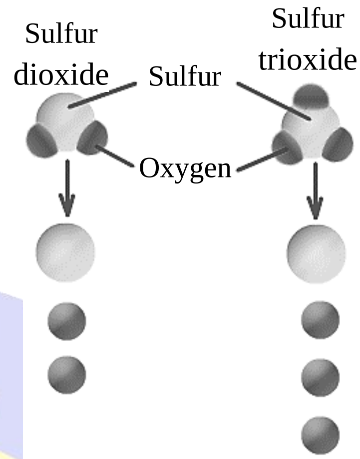

John Dalton (1776–1884)

'When two elements combine to form two or more compounds, in which fixed weight of one element combines with different weights of the other, will be in a simple numerical ratio'.

Ex: The weight of Oxygen that will combine with 12 g of carbon in CO and CO_2 is in the ratio of 1: 2



Ex: The weight of Oxygen that will combine with 32g of sulphur in SO_2 and SO_3 is the ratio of 2: 3

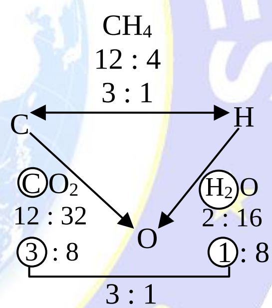


(d) LAW OF RECIPROCAL PROPORTION:

It is given by Richter.

The ratio of the weights of two elements A and B which combine separately with a fixed weight of the third element C is either the same or simple ratio of the weights in which A and B combine directly with each other.

Ex:



Ex: Ammonia contains 82.35% of nitrogen and 17.65% of hydrogen. Water contains 88.90% of oxygen and 11.10% of hydrogen. Nitrogen trioxide contains 63.15% of oxygen and 36.85% of nitrogen. Show that these data illustrate the law of reciprocal proportions.

Sol: In NH_3, 17.65g of H combine with N = 82.35g

1 g of H combine with N = $\frac{82.35}{17.65}$ g = 4.67g

In H_2O, 11.10 g of H combine with O = 88.90 g

1 g of H combine with O = $\frac{88.90}{11.10}$ g = 8.01g

Ratio of the weights of N and O which combine with fixed weight (=1g) of H = 4.67 : 8.01 = 1 : 1.7

In N_2O_3, ratio of weights of N and O which combine with each other =36.85 : 63.15 = 1 : 1.7

Thus, the two ratios are the same. Hence it illustrates the law of reciprocal proportions.

(e) Law of Gaseous Volume

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Under same conditions of temperature and pressure, whenever gases react together, the volumes of the reacting gases as well as products are in a simple whole number ratio.

Joseph Louis Gay Lussac

'at the same temperature and pressure, the volumes of gaseous reactants reacted and the volumes of gaseous products formed bear a simple ratio.'

Ex: $2H_2 + O_2 \rightarrow 2H_2O$ Volume ratio is 2: 1:2
2Vol 1Vol 2Vol

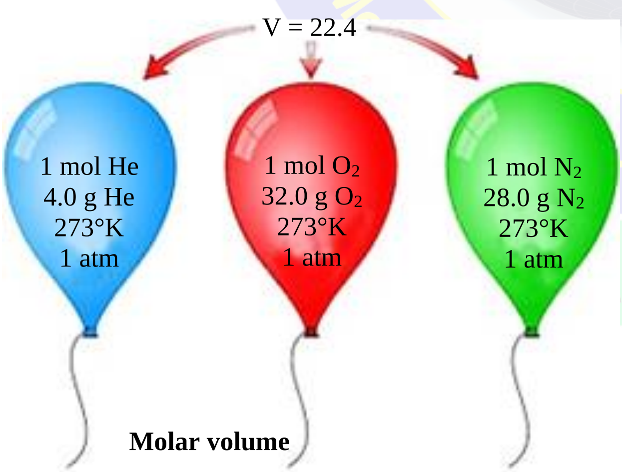
Ex: $N_2 + 3H_2 \rightarrow 2NH_3$ Volume ratio 1: 3:2
1Vol 3Vol 2Vol

(f) Avogadro's Law:

Amadeo Avogadro
(1776 – 1856)

It is given by "Amadeo Avogadro" :

Equal volumes of all gases under similar conditions of temperature and pressure contains equal number of molecules



Molar volume

STP (standard temperature and pressure) :
22.4 L for 1 mol of any gas.

Applications of Avogadro's law

(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

$$\text{Vapour density} = \frac{\text{Volume of definite amount of Gas}}{\text{Volume of same amount of Hydrogen}}$$

Or Vapour density

$$= \frac{\text{Weight of n molecules of Gas}}{\text{Weight of n molecules of Hydrogen}}$$

or Vapour density

$$= \frac{\text{Weight of one molecule of Gas}}{\text{Weight of one atom of hydrogen}\times 2}$$

or Vapour density

$$= \frac{\text{Molecular weight}}{2}$$

QUICK FOLLOW UP

Laws of Chemical Combination

Avagadro's Law

A constant pressure & temperature.
Volume is directly proportional to number of moles

Law of Conservation of mass

Matter can neither be created nor destroyed

Stoichiometry

Balance a Chemical Equation

Write correct formulas of reactant & products

↓

Balance Number of C atoms

↓

Balance Number of H atoms

↓

Balance Number of O atoms

↓

Verify Number of atoms of elements

Concentration term

$M = \dfrac{\text{No. of moles of solute}}{\text{Volume of solution in L}}$ $\quad m = \dfrac{\text{No. of moles of Solute}}{\text{Weight of solvent in kg}}$

$\dfrac{w}{v}\% = \dfrac{\text{Weight of solute kg}}{\text{Volume of solution in L}} \times 100$ $\quad \text{M.F.} = \dfrac{\text{Moles of Solute or solvent}}{\text{Total moles of solution}}$

$\dfrac{v}{v}\% = \dfrac{\text{Volume of solute in L}}{\text{Volume of solution in L}} \times 100$

$\dfrac{w}{w}\% = \dfrac{\text{Weight of solvent kg}}{\text{Weight of solution in kg}} \times 100$

$N = \text{Molarity} \times \text{n-factor}$ $\quad \text{Ppm} = \dfrac{\text{Weight of solute in kg}}{\text{Weight of solution in kg}} \times 10^6$

Limiting Reagent

The reactant that is entirely consume in a chemical reaction.

Dimensional Analysis

- Req. Unit =
  Given value × conservation factor
- Some useful conversation factors :

Length – 1 Å = 10^{-8} cm = 10^{-10} m
1 nm = 10^{-9} m, 1 pm = 10^{-12} m

Volume–1 L = 1000 mL
= 1000 cm^3 = 1 dm^3 = 10^{-3} m^3

Pressure – 1 atm 760 mm or torr = 101325 Pa
1 bar = 10^5 Nm^{-2} = 10^5Pa

Energy – 1 calorie = 4.184 J
1 eV = 1.6022×10^{-19} J
1 J = 10^7 ergs

Reaction in Solutions

$n = \dfrac{N}{N_A}, \; n = \dfrac{m}{M}, \; n = \dfrac{V}{22.4\text{ L}}$

Mass percent (%) = $\dfrac{w_{solute}}{w_{solution}} \times 100$

Mole Fraction, $x_A = \dfrac{n_A}{n_A + n_B}$,

$x_B = \dfrac{n_B}{n_A + n_B}$

Laws of Chemical Combination

Law of Definite Proportions

A given compound always contain same elements in the exact same proportions by mass

Gay Lussac's Law

A constant volume, pressure is directly proportional to temperature

Law of Multiple Proportions

If two elements can combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element are in ratio of small whole number

SOLVED EXAMPLES

Ex:1 One volume of hydrogen combines with one volume of chlorine to produce 2 volumes of hydrogen chloride.

Simple ratio = 1 : 1 : 2.

$$\boxed{1 \quad\quad 1 \quad\quad 2} \longleftarrow \text{Stoichiometry}$$

$$H_{2(g)} + Cl_{2(g)} \longrightarrow 2HCl_{(g)}$$

1Volume + 1 Volume 2 Volume

Ex:2 One volume of nitrogen combines with 3 volumes of hydrogen to from 2 volumes of ammonia.

Simple ratio = 1 : 3 : 2

$$\boxed{1 \quad\quad 3 \quad\quad 2}$$

$$N_{2(g)} + 3H_{2(g)} \longrightarrow 2NH_{3(g)}$$

1Volume + 3 Volume 2 Volume

Special Note: This law is used only for gaseous reaction. It relates volume to mole or molecules. But not relate with mass.

Ex:3 What is the weight of 3.01×10^{23} molecules of ammonia -

(1) 17 g (2) 8.5 g (3) 34 g (4) None

Sol: **(2)**

$\because$ 6.023×10^{23} molecules of NH_3 has weight = 17 g

$\therefore$ 3.01×10^{23} molecules of NH_3 has weight

$$= \frac{17 \times 3.01 \times 10^{23}}{6.023 \times 10^{23}} = 8.50 \text{ g}$$

Ex:4 At NTP, density of any gas has 0.00445 g/mL. Calculate the vapour density and molecular weight of the gas.

(1) 10, 70 (2) 20, 40 (3) 50, 100 (4) 30, 80

Sol: **(3)**

$$\text{V.D} = \frac{\text{Density of gas}}{\text{Density of } H_2} = \frac{0.004450}{0.000089} = 50$$

Molecular weight = 2 × V.D. = 2 × 50 = 100

Ex:5 How many molecules are present in one mL of water vapours at STP-

(1) 1.69×10^{19} (2) 2.69×10^{-19}

(3) 1.69×10^{-19} (4) 2.69×10^{19}

Sol: **(4)**

$\because$ 22.4 litre water vapour at STP has = 6.023×10^{23} molecules

$\therefore$ 1×10^{-3} litre water vapours at STP has

$$= \frac{6.023 \times 10^{23}}{22.4} \times 10^{-3} = 2.69 \times 10^{+19}$$

Ex:6 How many years it would take to spend Avogadro's number of rupees at the rate of 1 million rupees in one second –

(1) 19.098×10^{19} years (2) 19.098 years

(3) 19.098×10^{9} years (4) None of these

Sol: **(3)**

$\because$ 10^6 rupees are spent in 1sec.

$\therefore$ 6.023×10^{23} rupees are spent in

$$= \frac{1 \times 6.023 \times 10^{23}}{10^6} \text{ sec} = \frac{1 \times 6.023 \times 10^{23}}{10^6 \times 60 \times 60 \times 24 \times 365}$$

years = 19.098×10^{9} year

Ex:7 An atom of an element weighs 6.644×10^{-23} g. Calculate g atoms of element in 40 kg-

(1) 10 g atom (2) 100 g atom

(3) 1000 g atom (4) 10^4g atom

Sol: **(3)**

$\because$ weight of 1 atom of element = 6.644×10^{-23}g

$\therefore$ weight of 'N_A' atoms of element = $6.644 \times 10^{-23} \times 6.023 \times 10^{23}$ = 40 g

$\because$ 40 g of element has 1 g atom.

$\therefore$ 40 x 10^3g of element has $\dfrac{40 \times 10^3}{40}$

= 10^3g atom.

Ex:8 Calculate the number of Cl^- and Ca^{+2} ions in 222 g anhydrous $CaCl_2$ –

(1) $2N_A$ ions of Ca^{+2} 4 N_A ions of Cl^-

(2) $2N_A$ ions of Cl^- & $4N_A$ ions of Ca^{+2}

(3) $1N_A$ ions of Ca^{+2} & $1N_A$ ions of Cl^-

(4) None of these.

Sol: **(1)**

$\because$ mol. wt. of $CaCl_2$ = 111 g

$\because$ 111 g $CaCl_2$ has = N_A ions of Ca^{+2}

$\therefore$ 222g of $CaCl_2$ has $\dfrac{N \times 222}{111}$

= $2N_A$ ions of Ca^{+2} Also

$\because$ 111 g $CaCl_2$ has = 2N ions of Cl^-

$\therefore$ 222 g $CaCl_2$ has = $\dfrac{2N \times 222}{111}$ ions of Cl^-

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$= 4N_A$ ions of Cl^-.

Ex:9 The density of O_2 at NTP is 1.429g / litre. Calculate the standard molar volume of gas-

(1) 22.4 lit. (2) 11.2 lit
(3) 33.6 lit (4) 5.6 lit.

Sol: **(1)**

$\because$ 1.429 g of O_2 gas occupies volume = 1 litre.

$\therefore$ 32g of O_2 gas occupies = $\frac{32}{1.429}$

= 22.4 litre/mol.

Ex:10 Which of the following will weigh maximum amount-

(1) 40 g iron
(2) 1.2 g atom of N
(3) 1×10^{23} atoms of carbon
(4) 1.12 litre of O_2 at STP

Sol: **(1)**

(1) Mass of iron = 40 g

(2) Mass of 1.2 g atom of N = 14 × 1.2 = 16.8 g

(3) Mass of 1×10^{23} atoms of C = $\frac{12\times1\times10^{23}}{6.023\times10^{23}}$

= 1.99 g.

(4) Mass of 1.12 litre of O_2 at STP = $\frac{32\times1.2}{22.4}$

= 1.6 g

Ex:11 How many moles of potassium chlorate to be heated to produce 11.2 litre oxygen -

(1) $\frac{1}{2}$ mol (2) $\frac{1}{3}$ mol
(3) $\frac{1}{4}$ mol (4) $\frac{2}{3}$ mol.

Sol: **(2)**

$$2\,KClO_3 \rightarrow 2KCl + 3O_2$$

Mole for reaction 2 2 3

$\because$ 3 × 22.4 litre O_2 is formed by 2 mol $KClO_3$

$\therefore$ 11.2 litre O_2 is formed by $\frac{2\times11.2}{3\times22.4}$

= $\frac{1}{3}$ mol $KClO_3$

Ex:12 Calculate the weight of lime (CaO) obtained by heating 200 kg of 95% pure lime stone ($CaCO_3$).

(1) 104.4 kg (2) 105.4 kg
(3) 212.8 kg (4) 106.4 kg

Sol: **(4)**

$\because$ 100 kg impure sample has pure $CaCO_3$ = 95 kg

$\therefore$ 200 kg impure sample has pure $CaCO_3$

= $\frac{95\times200}{100}$ = 190 kg.

$$CaCO_3 \rightarrow CaO + CO_2$$

Mol. Wt. 100 56 44

$\because$ 100 kg $CaCO_3$ gives CaO = 56 kg.

$\therefore$ 190 kg $CaCO_3$ gives CaO = $\frac{56\times190}{100}$ = 106.4 kg.

Ex.:13 The chloride of a metal has the formula MCl_3. The formula of its phosphate will be-

(1) M_2PO_4 (2) MPO_4
(3) M_3PO_4 (4) $M(PO_4)_2$

Sol: **(2)**

$AlCl_3$ as it is $AlPO_4$

Ex:14 A silver coin weighing 11.34 g was dissolved in nitric acid. When sodium chloride was added to the solution all the silver (present as $AgNO_3$) was precipitated as silver chloride. The weight of the precipitated silver chloride was 14.35 g. Calculate the percentage of silver in the coin –

(1) 4.8 % (2) 95.2% (3) 90 % (4) 80%

Sol: **(2)**

$Ag + 2HNO_3 \rightarrow AgNO_3 + NO_2 + H_2O$
108

$AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$
143.5

$\therefore$ 143.5 g of silver chloride would be precipitated by 108 g of silver.

or 14.35 g of silver chloride would be precipitated 10.8 g of silver.

$\because$ 11.34 g of silver coin contain 10.8 g of pure silver.

$\therefore$ 100 g of silver coin contain $\frac{10.8}{11.34} \times 100$

= 95.2 %.

Ex:15 Calculate how many methane molecules and how many hydrogen and carbon atoms are there in 25.0 g of methane?

Sol: moles of CH_4 = $\frac{25}{16}$

$\therefore$ No. of CH_4 molecules = $\frac{25}{16} \times 6.02 \times 10^{23}$

$= 9.41 \times 10^{23}$

$\because$ 1 molecule of CH_4 contains one carbon atom and four hydrogen atom

$\therefore$ No. of C atom = 9.41×10^{23}

$\therefore$ No. of H atoms = $4 \times 9.41 \times 10^{23} = 37.64 \times 10^{23}$

Ex:16 The vapour density of a mixture containing NO_2 and N_2O_4 is 38.3 at 27^0C. Calculate the mole of NO_2 in 100 mole mixture.

Sol: Molecular wt. of mixture of NO_2 and N_2O_4
= 38.3 × 2 = 76.6
Let x moles of NO_2 are present in 100 mol mixture
∴ moles of N_2O_4 = (100 – x)
wt. of NO_2 + wt. of N_2O_4 = total wt. of mixture
(x × 46) + (100 – x) × 92 = 100 ×76.6 g
⇒ x = 33.48 moles

Ex:17 The percentage by volume of C_3H_8 in a mixture of C_3H_8, CH_4 and CO is 36.5. Calculate the volume of CO_2 produced when 100 mL of the mixture is burnt in excess of O_2.

Sol: Let a, b and c ml be volumes of C_3H_8, CH_4 and CO respectively in 100 ml given sample than
a + b + c = 100 and a = 36.5 ml
Now CO_2 is formed as a result of combustion of mixture as follows –

(1) $C_3H_8 + 5O_2 \longrightarrow 3CO_2 + 4H_2O$
a vol. 3a vol.

(2) $CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$
b vol. b vol.

(3) $CO + 1/2\ O_2 \longrightarrow CO_2$
c vol. c vol.

Total vol. of CO_2 produced = 3a + b + c = 2a + (a + b + c) = 2 × 36.5 + 100 = 173 ml.

Ex:18 1.0 g of metal nitrate gave 0.86 g of metal sulphate. Calculate equivalent wt. of metal.

Sol: $M(NO_3)_n \longrightarrow M_2(SO_4)_n$ (n = valency of metal)
g eq. $M(NO_3)_n$ = g eq. of $M_2(SO_4)_n$

$$\frac{1.0}{E(M)+E(NO_3^-)} = \frac{0.86}{E(M)+E(SO_4^{2-})}$$

$$\Rightarrow \frac{1}{E+\frac{62}{1}} = \frac{0.86}{E+\frac{96}{2}} \Rightarrow E = 38$$

Ex:19 2 g of metal in H_2SO_4 gives 4.51 g of the metal sulphate. The specific heat capacity of metal is 0.057 cal/g. Calculate the valency and atomic weight of metal.

Sol: g eq. of metal = g eq. of metal sulphate.

$$\frac{2}{E} = \frac{4.51}{E+E(SO_4^{2-})} \text{ (E = eq. wt of metal}$$

$$\frac{2}{E} = \frac{4.51}{(E+48)} \Rightarrow E = 38.24$$

At. wt. × sp heat ≃ 6.4

approx. At.wt. = $\frac{6.4}{0.057} = 112.28$

∵ Valency of metal = $\frac{\text{At. wt.}}{\text{Eq. wt.}} = \frac{112.28}{38.24} = 2.93$

= 3 (∵ Valency is integer)

∴ Exact at. wt. of metal = Eq. wt × Valency
= 38.24 × 3 = 114.72

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