# Mole Concept

## CHEMISTRY

- Chemistry is defined as the study of the composition, properties and interaction of matter.

### Branches of Chemistry

**(1) Physical Chemistry**

The discipline of chemistry concerned with the way in which physical properties of substances depend on and influence their chemical structure, properties, and reactions.

**(2) Inorganic Chemistry**

The discipline of chemistry in which structure, composition, and behavior of inorganic compounds. All the substances other than the carbon-hydrogen compounds are classified under the group of inorganic substances.

**(3) Organic Chemistry**

The branch in which the study of the structure, composition and the chemical properties of organic compounds is known as organic chemistry.

**(4) Biochemistry**

The discipline which deals with the structure and behavior of the components of cells and the chemical processes in living beings is known as biochemistry

**(5) Analytical Chemistry**

The branch of chemistry dealing with separation, identification and quantitative determination of the compositions of different substances

## MATTER

- It exists in three physical states, e.g., Solid, Liquid and Gas. It also has two other states namely Bose-Eistein condensite and Plasma

**Concept Ladder**

Chemistry, from the ancient Egyptian word "khēmia" meaning transmutation of earth, is the science of matter at the atomic to molecular scale, dealing primarily with collections of atoms, such as molecules, crystals, and metals.

**Rack your Brain**

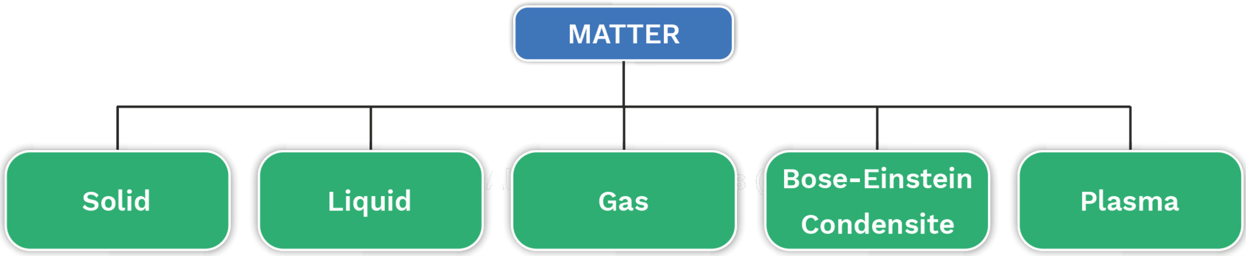
How was mass of single carbon aotm determined?

**Concept Ladder**

Antoine Lavoisier is known as father of chemistry. He developed an experimentally based theory of the chemical reactivity of oxygen and coauthored the modern system for naming chemical substances.

**Definition**

**Matter** is anything that occupies mass and space.




### Characteristics

(1) Solids have definite volume and definite shape.

(2) Liquids have definite volume but definite shape. They take the shape of the container in which they are kept.

(3) Gases ahve neither definite volume nor definite shape. They occupy completelty the container in which they are kept.

These three states are interconvertible on changing the conditions of temperature and pressure.

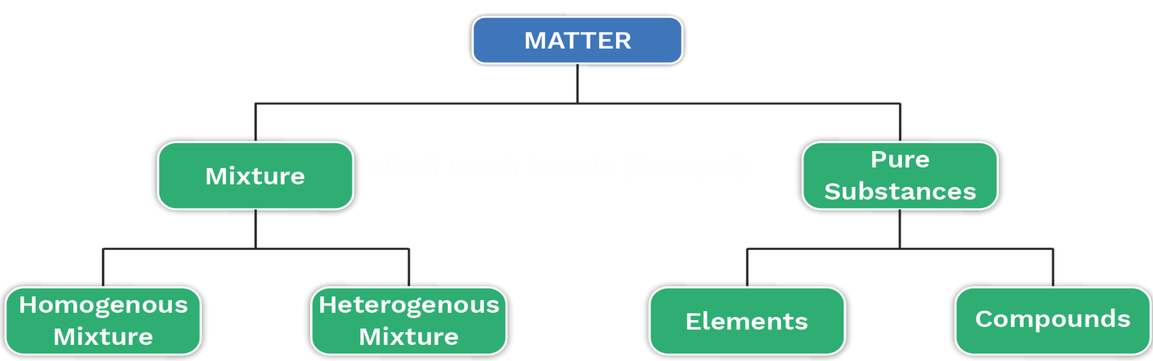
$$\text{Solid} \underset{\text{Cool}}{\overset{\text{Heat}}{\rightleftharpoons}} \text{Liquid} \underset{\text{Cool}}{\overset{\text{Heat}}{\rightleftharpoons}} \text{Gas}$$

**Concept Ladder**

In chemistry, a substance is a form of matter that has constant chemical composition and characteristic properties. It cannot be separated into components by physical separation methods, i.e. without breaking chemical bonds. They can be solids, liquids or gases.

### Classification of Matter On the Basis of Purity

Matter can be classified broadly as mixture or pure substances, which can be further subdivided as shows below.



**(1) Mixture**

Generally pure substances are added together to form a mixture. Also, a mixture can be obtained by mixing two mixtures. For example, sugar solution in water, air, tea etc.

**(2) Homogenous Mixture**

A mixture in which the components completely mix with each other and its composition is uniform throughout. For example, salt solution, sugar solution, air etc.

**(3) Heterogenous Mixture**

A mixture which doesn’t have same composition throughout and different components sometimes can be observed. For example, grains and pulses along with some dirt (often stone) pieces , mixture of salt and sugar etc.

**(4) Pure Substances**

They have fixed composition, whereas mixture may contain the components in any ratio and its composition is variable. For example, gold, silver, copper, water, glucose etc.

**(5) Element**

Element has only one type of particles, atoms or molecules. For example silver, copper, sodium, hydrogen, oxygen etc.

**(6) Molecule**

When two or more atoms combine molecule is formed. Hydrogen, oxygen, and nitrogen gases consists of molecules in which atoms of same elements combine to give their respective molecules.

**(7) Compound**

Compounds are always formed when substances combine in different ratios by mass. For example water ammonia, carbon

**Definition**

A **mixture** contains two or more substances present in it in any ratio which are called its components.

**Concept Ladder**

Alloys are mixtures of two or more metals or a metal and a non-metal and cannot be separated into their components by physical methods. For example, brass is a mixture of approximately 30% zinc and 70% copper.

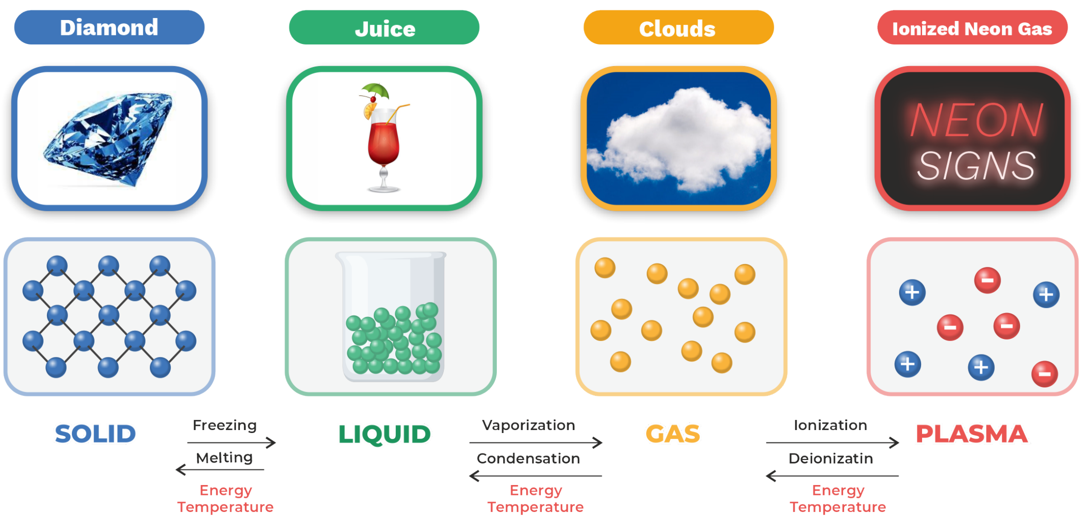
**Rack your Brain**

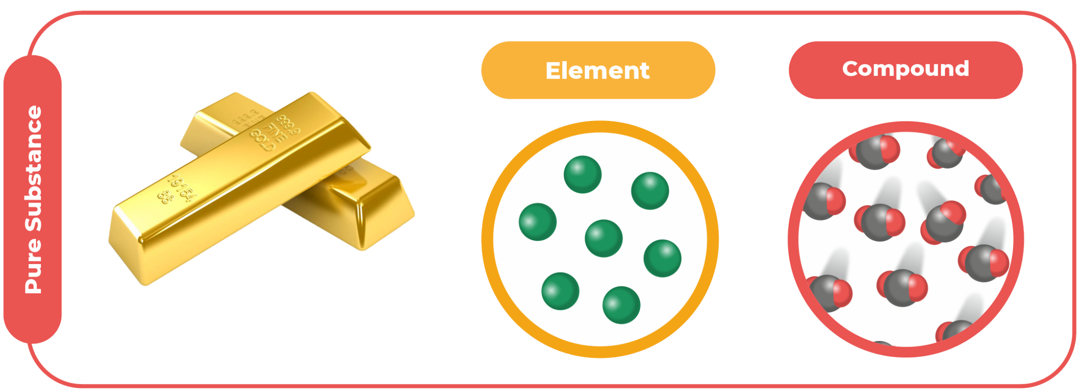
How do we judge whether milk, ghee, butter, salt, spices, mineral water or juice that we buy from the market are pure?

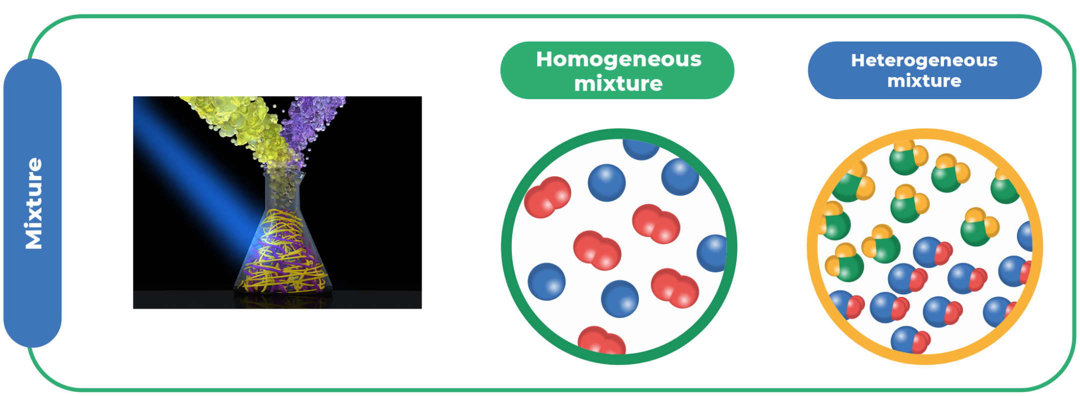
**Definition**

When two or more atoms of different elements combine, the molecule of a compound is formed.

# STATES OF MATTER



monoxide, sodium chloride etc.

### Properties of Matter and Their Measurements

- The properties of a substance have unique characteristics and are classified into physical and chemical properties.

**(1) Physical Properties**

Those properties which can be measured or observed without changing the identity or composition of the substance are known as physical properties, .e.g, colour, melting point, boiling point, odour etc.

**(2) Chemical Properties**

Those properties which describe a matter's 'potential' to undergo some chemical changes are konwn as chemical properties, e.g., characteristics reactions of different substances which include acidity or basicity, combustibility etc.

**(3) Measurement**

Any quantitative observation represented by a number followed by a unit in which it is measued is called measurement, such as length, are, volume etc.

### PHYSICAL QUANTITIES AND SI UNITS OF MEASUREMENT

- $11^{th}$ general conference of weights and measures recommended the use of international system of units in 1960. SI Units is abbreviated as (after the French expression La System International de units).

### Fundamental Units

- The SI system has seven basic/fundamental units of physical quantities as follows:

**Concept Ladder**

Physical properties like melting and boiling points can be the result of the components present inside a system.

**Rack your Brain**

How do we judge whether milk, ghee, butter, salt, spices, mineral water or juice that we buy from the market are pure?

**Concept Ladder**

There are other unit systems used for measurement of different physical and chemical quantities For example, FPS (Foot, poundal, second), CGS (centimeter, gram, second).

**Rack your Brain**

Can you guess what is BTU?

| Physical Quantity | Abbreviation | Name of unit | Symbol |
|---|---|---|---|
| time | t | second | s |
| mass | m | kilogram | kg |
| length | l | metre | m |
| temperature | T | kelvin | K |
| electric current | I | ampere | A |
| light intensity | lv | candela | Cd |
| amount of substance | n | mole | mol |

**Concept Ladder**

Fundamental units are the basis of derived units.. Derived units are obtained when fundamental units are divided or multiplied together but cannot be obtained on addition or substraction.

### Derived Units

- The units obtained by combination of basic units are known as derived units e.g. velocity is expressed as distance/time. Hence unit is m/s or ms–1. Some common derived units are:

| Physical Quantity | Definition | SI Unit |
|---|---|---|
| volume | length cube | $m^3$ |
| area | length square | $m^2$ |
| speed | distance travelled | $ms^{-1}$ per unit time |
| acceleration | speed changed | $ms^{-2}$ per unit time |

## Standard Prefixes

| Fraction | Prefix | Symbol | Multiple | Prefix | Symbol |
|---|---|---|---|---|---|
| $10^{-1}$ | deci | d | $10^{1}$ | Deka | da |
| $10^{-2}$ | centi | c | $10^{2}$ | Hecta | h |
| $10^{-3}$ | milli | m | $10^{3}$ | kilo | k |
| $10^{-6}$ | micro | m | $10^{6}$ | Mega | M |
| $10^{-9}$ | nano | n | $10^{9}$ | Giga | G |
| $10^{-12}$ | pico | p | $10^{12}$ | Tera | T |
| $10^{-15}$ | femto | f | $10^{15}$ | Peta | P |
| $10^{-18}$ | atto | a | $10^{18}$ | Exa | E |

## Precision and Accuracy

**Precision**

Precision is the closeness of various measurements done for the same quantity

**Accuracy**

Accuracy is the agreement of a subjected value to the true value.

**Ex:** Let the true weight of a substance be 3.00g. The measurement reported by three students are as follows:

| Sudent | Measurements/g | | Average/g |
|---|---|---|---|
| | 1 | 2 | |
| A | 2.95 | 2.93 | 2.94 |
| B | 3.01 | 2.99 | 3 |
| C | 2.94 | 3.05 | 2.99 |

POOR ACCURACY
POOR PRECISION

**Concept Ladder**

The accuracy of any measurment depends upon two parameters. First being the device and second being the skill of operator.

**Case of A student**

It is the case of precision but no accuracy since measurements one close but not accurate.

**Case of B student**

Measurements are very close (precision) and accurate (Accuracy).

**Case of C student**

Measurement is not close (no precision) and not accurate (no accuracy).

GOOD ACCURACY
GOOD PRECISION

## SOME IMPORTANT DEFINITION

### Mass and Weight

- It is the quantity of matter present in it while weight is the force exerted by gravity on an object.
- The mass of a substance is constant. Weight varies from one place to another due to change in gravity.
- SI unit of mass is kg.

### Volume

- Volume is often quantified numerically using the SI derived unit, the cubic meter ($m^3$).
- Volume of liquids or solutions is measured by using burette, pipette, graduated cylinder, or volumetric flask.

### Density

- The mass occupied per unit volume of a substance is known as Density. The symbol most often used for density is $\rho$
- SI unit of density is kg $m^{-3}$.
- Other units of density include lb $ft^{-3}$ and g $cm^{-3}$.
- Density is usually calculated with respect to a standard substance. This is known as relative density.


**Definition**

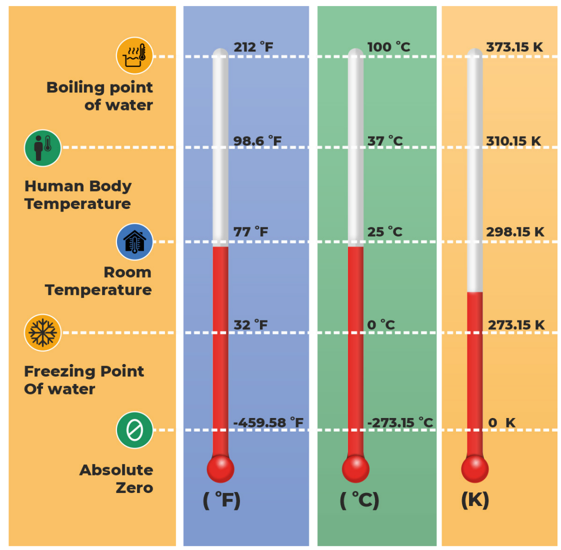
Volume is the quantity of three-dimensional space enclosed by some closed boundary, for example, the space that a substance (solid, liquid, gas, or plasma) or shape occupies or contains.

## Temperature

- Temperature is a physical property of matter which tell us about the degree of heat content.
- There are three common scales for measurement of temperature — K (kelvin, °F (degree Fahrenheit) and)°C (degree Celsius).
- The temperature on two scales is related to each other by the following relationship:

  °F = 9/5(°C) + 32

  K = °C + 273-15




## LAWS OF CHEMICAL COMBINATION

- The combination of elements to form compounds chemically is the result of the five basic laws:

**Laws of chemical combination**
- Laws of Conversion of Mass
- Laws of Definite Proportions
- Laws of Multiple Proportions
- Gay Lussac's Law of Gaseous Volumes
- Avogadro Law

### Concept Ladder

Relation between different temperature scales.

$$\frac{T_{°C}}{100} = \frac{T_{°F} - 32}{180} = \frac{T_K - 273.15}{100}$$

## Law of Conservation of Mass

- The law states that matter can neither be created nor destroyed.
- This law was given by Antoine Lavoisier in 1789.

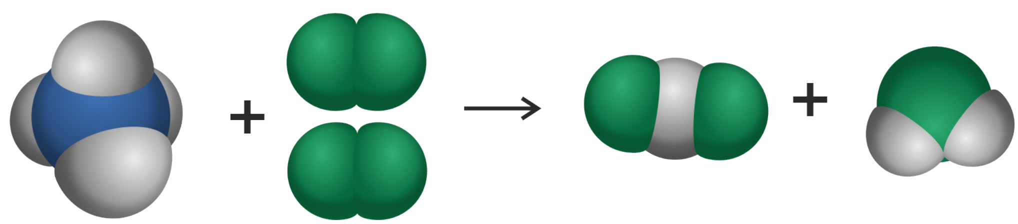
### Rack your Brain

Is law of conservation of mass valid when particles of light break from the packet of quanta?

- He performed experimental studies for combustion reactions for reaching to the above conclusion.
- This law did form the basis for several later developments in chemistry. In fact, this was the result of absolutely exact measurement of masses of reactants and products, and carefully planned experiments done by Lavoisier.

**Rack your Brain**

What will happen when hydrogen and sulphur combine in the ratio 1:16 by mass?



$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$$

**Q1** Five grams of $KClO_3$ yid 3.04 1 g and 1.36 L of oxygen at standard temperature and pressure. Show that these figures support the law of conservation of mass within limits of ±0.4% error.

**A1** According to gram-molecular volume law, 22.4 L of all gases and vapours at STP weight equal to their molecular weights denoted in grams.

$\therefore$ Weight of 1.36 L of oxygen at STP $= \dfrac{32 \times 1.36}{22.4} = 1.943$ g

Weight of KCl formed = 3.041 g (given)

$\therefore$ Total weight of product ($KCl + O_2$) = 3.041 + 1.943 = 4.984 g

Error = 5 – 4.984 = 0.001 g

$\therefore$ % error $= \dfrac{0.016 \times 100}{5} = 0.32$

Hence, the law of conservation of mass is valid within limits of –0.4% error. Thus, the law is supported.

## Law of Definite Proportions

- This law was given by French chemist, Joseph Proust.
- It was stated by him that a given compound always contains the same proportion of elements by weight.
- Proust considered two samples of cupric carbonate — one of the samples from natural origin and the other was of synthetic.
- He found that the composition of elements present in it was similar for both the samples as shown below:

| | % of copper | % of oxygen | % of carbon |
|---|---|---|---|
| Natural sample | 51.35 | 9.74 | 38.91 |
| Synthetic sample | 51.35 | 9.74 | 38.91 |

- Thus, irrespective of the source, a given compound always contains same elements in the same proportion. The validity of this law has been confirmed by various experiments. It is sometimes also referred to as Law of definite composition.

**Concept Ladder**

Law of constant composition is not true for all types of compounds but true only for the compounds obtained from one isotope.

**Rack your Brain**

Can we apply the law of definite proportion in case of non-stoichiometric compounds and polymers?

**Q2** 0.7 g of iron reacts directly with 0.4 g of sulphur to form ferrous sulphide. If 2.8 g of iron is dissolved in dilute HCl and excess of sodium sulphide solution is added, 4.4 g of iron sulphide is precipitated. Prove the law of constant composition.

**A2** The ration of the weight of iron and sulphur in the first sample of the compound is Fe : S :: 0.7 : 0.4 or 7 : 4.
According to the second experiment, 2.8 g of iron gives 4.4 g ferrous sulphide, or 2.8 g Fe combines with S = 4.4 – 2.8 = 1.6 g
Therefore, the ratio of the weights of Fe : S :: 2.8 : 1.6 or 7.8.
Since the ratio of the weights of the two elements is same in both the cases, the law of constant composition is true.

## Law of Multiple Proportions

- The law was proposed by Dalton in 1803.
- According to this law, when two elements combine to form more than one compound, the mass of one element that combines with fixed mass of the other element, are in the ratio of simplest whole numbers.
- For example, $H_2$ combines with $O_2$ to form two compounds, namely, water and $H_2O_2$.

Hydrogen + Oxygen $\longrightarrow$ Water
2g 16g 18g

Hydrogen + Oxygen $\longrightarrow$ Hydrogen Peroxide
2g 32g 34g

**Concept Ladder**

The law, which was based on Dalton's observations of the reactions of atmospheric gases, states that when elements form compounds, the proportions of the elements in those chemical compounds can be expressed in small whole number ratios.

**Q3** Elements X and Y form two different compounds. In the first compound, 0.324 g X is combined with 0.471 g Y. In the second compounds, 0.117 g X is combined with 0.509 g Y. Show that these data illustrate the law of multiple proportions.

**A3** In the first compound 0.324 g of X combines with 0.471 g of Y. In the second compound 0.117 g of X combines with 0.509 g of Y.

Therefore, 0.324 g of X combines with the weight of Y $= \dfrac{0.509 \times 0.324}{0.117} = 1.4095$ g

Now, the weights of Y that combine with the same weight of X, i.e., 0.324 g of it, are in the ratio of 0.471 : 1.4095 or 1 : 3. The ratio, being simple, illustrates the law of multiple proportions.

## Gay Lussac's Law of Gaseous Volumes

- This law was given in 1808 by Gay Lussac.
- It was observed that when gases combine together in a chemical reaction they combine in a simple ratio by volume provided all the gases are at same pressure and temperature.
- 1000 mL of $H_2$ combine with 500 mL of $O_2$ to give 1000 mL of $H_2O$ vapour.

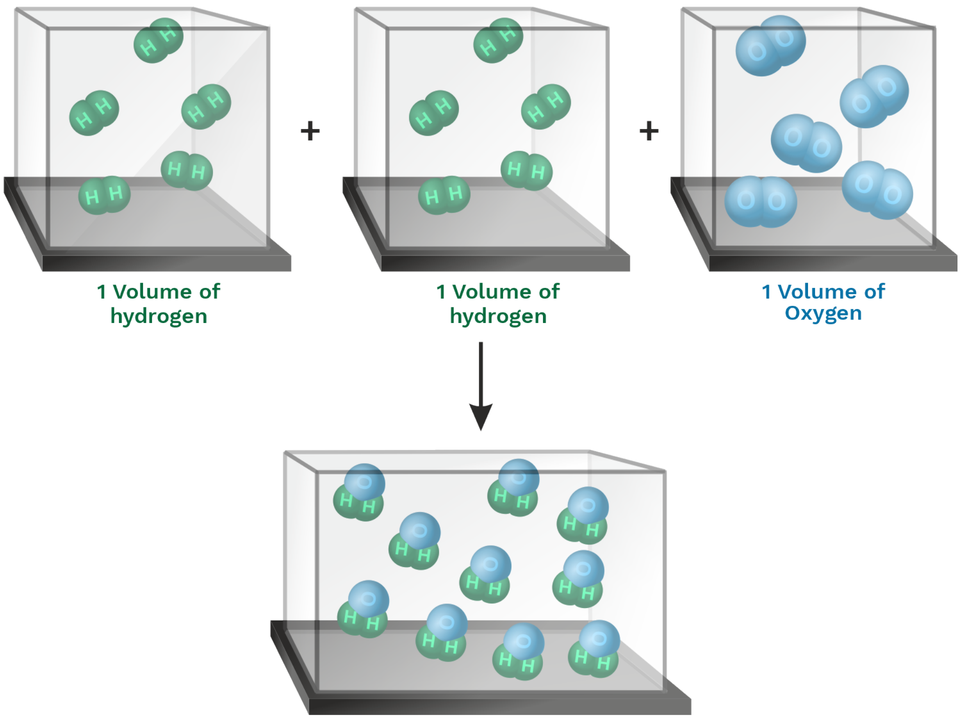
$H_2$ + $O_2$ $\longrightarrow$ $H_2O$
1000 mL 500 mL 1000 mL

**Previous Year's Question**

Equal masses of $H_2$, $O_2$ and methane have been taken in a container of volumes V at temperature 27°C in identical conditions. The ratio of the volumes of gases $H_2$:$O_2$:methane would be

**[AIPMT-2014]**

(1) 8 : 16 : 1 (2) 16 : 8 : 1
(3) 16 : 1 : 2 (4) 8 : 1 : 2




**Q4** Air contains 21% oxygen by volume. Calculate the theroretical volume of air which will be required for burning completely 500 cubic ft of acetylene gas ($C_2H_2$).

**A4** $$\underset{2\ \text{vol}}{2C_2H_2} + \underset{5\ \text{vol}}{5O_2} \rightarrow \underset{4\ \text{vol}}{4CO_2} + \underset{2\ \text{vol}}{2H_2O}(\text{steam})$$

According to the above equation:

2 volumes of acetylene require 5 volumes of $O_2$ for combustion.

$\therefore$ 500 eu. ft. of acetylene will require $O_2 = \dfrac{5 \times 500}{2} = 1250$ cu. ft.

Hence, the quantity of air required $= \dfrac{100 \times 1250}{21} = 5952$ cu. ft.

### Avogadro law

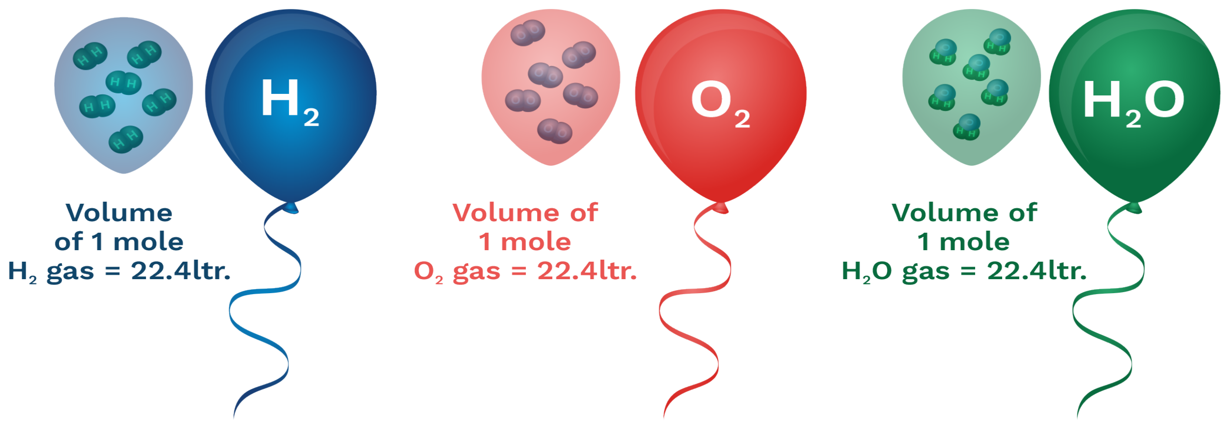
- Avogadro proposed that when equal volumes of gases at same temperature and pressure they should contain equal number of molecules.

- Avogadro gave the distinction between atoms and molecules which is quite understandable in the present times.
- If we consider, reaction of $H_2$ and $O_2$ to produce $H_2O$, we see that two volumes of $H_2$ combine with one volume of $O_2$ to give two volumes of $H_2O$ without leaving any unreacted $O_2$.

**Rack your Brain**

A container of 5 liter consists $CO_2$ gas. Another container of half the volume consists $O_2$ at same condition. Would it follow Avogadro's law?

## DALTON'S THEORY OF ATOM




J. Dalton gave his famous theory of atom in 1803. The main postulates of this theory

- Atom is considered as a hard, dense and_ the smallest indivisible particle of matter.
- Each element consists of a identical kind of atoms.
- The properties of elements differ because of difference in the kinds of atoms contained in them.
- This theory provides a satisfactory basis for the law of chemical combination.

### Limitations of Dalton's Theory

- This theory fails to explain why the atoms of different kinds should differ in mass, valency etc.

**Concept Ladder**

Dalton's atomic theory had certain explanation of Laws of Chemical Combination :

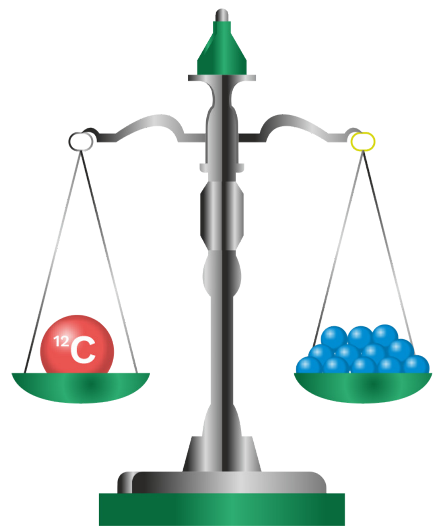
Law of conservation of mass : Matter consists of atoms and can neither be created nor destroyed.

Law of constant composition: When atoms of same or different elements combine together to form compounds, they combine in a fixed ratio, a simple whole number ratios.

- The discovery of isobars and isotopes demonstrated that atoms of same elements may have different atomic masses (isotopes) and atoms of different kinds may have same atomic masses (isobars).
- The discovery of various sub-aomic particles like electrons, protons, X-rays, etc. during the late 19th century lead to idea that atom was no longer an indivisible and the smallest particle of the matter.

**Rack your Brain**

Can you find law of multiple proportion and law of reciprocal proportion form Dalton's Atomic Theory Postulate?




### ATOM

- Each element is formed of smallest particles called 'ATOM'.
- Atom is derived from Greek language Atoms means 'Not to be cut'.

### Atomic Mass

- Atomiic mass of an element can be defined by a number which indicates how many times the mass of one atom of the element is heavier in comparison to $\frac{1}{12}$th part of the mass of one atom of Carbon-12.
- $$\text{Atomic mass} = \frac{[\text{Mass of an atom of the element}]}{\frac{1}{12}\times[\text{Mass of an atom of carbon-12}]} = \frac{\text{Mass of an atom in amu}}{1\text{ amu}}$$

### Atomic mass unit (amu) or Unified mass (u)

- The quantity $\frac{1}{12}$ × mass of an atom of C-12 is known as atomic mass unit.
- $\therefore 1\text{ a.m.u.} = \frac{1.9924\times10^{-26}}{12}\text{kg}$

**Definition**

Average atomic mass is defined as the average mass of isotopes of an element naturally present.

### Average Atomic Mass (AAM)

- The atomic mass as the weighted average mass of all naturally occurring isotopes of the element.
- $$\text{A.A.M.} = \left(\frac{\%\text{ abundance isotope 1}}{100}\right)\times(A_1) + \left(\frac{\%\text{ abundance isotope 2}}{100}\right)\times(A_2) + ......$$

**Q5** Chlorine is a mixture of two isotopes with atomic masses 35u and 37u and they are present in the ration of 3:1. Determine average atomic mass of chlorine?

**A5** Average atomic mass of chlorine

$$= \frac{35 \times 0.75 + 37 \times 0.25}{0.75 + 0.25} = 35.49 \text{ u}$$

**Concept Ladder**

Mass spectroscopy and osmotic pressure measurements are used for the determination of molecualr weight of any compound.

### Molecular Mass

- Molecular mass is the number which indicates how many times one molecule of a substance is heavier in comparison to $\frac{1}{12}$th of the mass of one aotm of C-12.
- $$\text{Molecular mass} = \frac{\text{Mass of one molecule of the substance (in amu)}}{\frac{1}{12} \times [\text{Mass of an atom of C-12}]}$$

$$= \frac{\text{Mass of one molecule of substance (in amu)}}{1 \text{ amu}}$$

**Q6** Calculate molecular mass of the following molecules:
(i) sulphuric acid ($H_2SO_4$) (ii) Glucose ($C_6H_{12}O_6$)

**A6** (i) Molecular mass of $H_2SO_4$
= 2(1.008 u) + 32.0 u + 4 (16.0) u
= 2.016 u + 32.0 u + 64.0 u = 98.016 u
(ii) Molecualr mass of $C_6H_{12}O_6$
= 6(12.011 u) + 12(1.008 u) + 6(1600 u)
= 72.066 u + 12.096 u + 96.00 u = 180.162 u

**Q7** Calculate molecular mass of glucose ($C_{12}H_{22}O_{11}$) molecule.

**A7** Molecular mass of glucose ($C_{12}H_{22}O_{11}$)
= 12(12.011 u) + 22(1.008 u) + 11(16.00 u)
= 342.308 u

### Average Molecular Mass (AMM)

- Average Molecular Mass

$$= \frac{\text{total mass}}{\text{total mole of molecules}}$$

- Let a sample contains $n_1$ mole of molecules with molecular mass $M_1$ and $n_2$ mole of molecules with molecualr mass $M_2$, then

$$M_{av} = \frac{n_1M_1 + n_2M_2}{n_1 + n_2}$$

### Formula Unit Mass

- The mass of a formula unit present in ionic compound is known as formula unit mass.
- For example, the formula unit mass of sodium chloride (NaCl) is 58.5 u.

### Gram Atomic Mass (GAM)

- Gram atomic mass of an atom expressed in grams.
- Number of gram atoms

$$= \frac{\text{Mass of element in grams}}{\text{GAM of element}}$$

### Gram Molecular Mass (GMM)

- That amount of substance having mass in grams which is equal to its molecular mass of a substance expressed in grams is known as gram molecular mass.
- It is also called one gram molecule.
- Number of gram molecules

$$= \frac{\text{Weight of substance}}{\text{GMM of substance}}$$

**Previous Year's Question**

An element, X has the followign isotopic composition :

$^{200}X$:90% $^{199}X$:8.0%

$^{202}X$:2.0%

**[AIPMT-2007]**

(1) 201 amu (2) 202 amu
(3) 199 amu (4) 200 amu

**Rack your Brain**

There is a compound $XCl_2$ which is ionic in nature. Can this compound have molecular mass?

**Definition**

The atomic mass of an element expressed in grams is called gram atomic mass.

**Definition**

The molecualr mass of a compound expressed in grams is called gram molecular mass.

## MOLE CONCEPT AND MOLAR MASSES

- A mole is amount of a substance which consists of as many entities as there are atoms in 12 g (or 0.012 kg) of the C-12 isotope.
- The mass of a C-12 was calculated by a mass spectrometer and was found to be equal to $1.992648 \times 10^{-23}$ g.
- Number of atoms in one mole of carbon

$$= \frac{12 \text{ g mol}^{-1} \text{ of } {}^{12}\text{C}}{1.992648 \times 10^{-23} \text{ g}\left({}^{12}\text{C atom}\right)^{-1}}$$

$= 6.0221367 \times 10^{23}$ atoms $mol^{-1}$

$\approx 6.022 \times 10^{23}$ atoms $mol^{-1}$

### Previous Year's Question

If Avogadro number $N_A$, is changed from $6.022 \times 10^{23}$ $mol^{-1}$ to $6.022 \times 10^{20}$ $mol^{-1}$, this would change

**[AIPMT-2015]**

(1) the mass of one mole of carbon
(2) the ratio of chemical species to each other in a balanced equation
(3) the ratio of chemical species to each other in a compound
(4) the definition of mass in units of grams.

### Molar Mass

- The numercal value of molar mass of a substance in grams is numerically equal to atomic or molecular formula mass in u. Units of molar mass are g $mol^{-1}$ or kg $mol^{-1}$.

$\therefore$ Molar mass of $CO_2$ = 12.011 + 2(16.0) = 44.011 g $mol^{-1}$

Molar mass of NaCl = 23.0 + 35.5 = 58.5 g $mol^{-1}$

### Definition

The mass of one mole of a substance in called its molar mass.

### Molar Volume

- According to Avogadro's hypothesis, equal volumes of different gases under similar conditions of temperature and pressure contian equal number of molecules.
- A mole of a gas at STP (standard temperature and pressure), viz., 1 atm and 273 K (0°C) contains NA molecules ($6.022 \times 10^{23}$).
- 1 mole of a gas at 1 bar and 273 K consiting of $N_A$ molecules will have a volume of 22.7 liters which is considered to be **new STP** condition.

### Concept Ladder

**Loschdmidt Number**

The number of molecules present in 1 $cm^3$ of an ideal gas at STO is called Loschdmidt number

$$= \frac{6.022 \times 10^{23}}{22400 \text{ mL}} = 2.69 \times 10^{19}$$

## BASIC CONCEPTS OF ATOM

### To calculate no. of p, n and $e^-$

**(a) In case of neutral atom**

| Atom | p | $e^-$ | n |
|---|---|---|---|
| Carbon (C) | 6 | 6 | 6 |
| Nitrogen (N) | 7 | 7 | 7 |

**(b) In case of ions**

| Ion | p | $e^-$ | n |
|---|---|---|---|
| Oxide ion ($O^{2-}$) | 8 | 8 + 2 = 10 | 8 |
| Nitride ion ($N^{3-}$) | 7 | 7 + 3 = 10 | 7 |

**(c) In case of molecule**

**Ex:** $CH_4$

Total number of atoms in $CH_4$ = 5

| Molecule | Element | p | $e^-$ | n |
|---|---|---|---|---|
| $CH_4$ | Carbon | 6 | 6 | 6 |
| | Hydrogen (1 × 4) | 4 | 4 | 0 |
| Total | | **10** | **10** | **6** |

**Ex:** $N_2$

Total number of atoms in $N_2$ = 2

| Molecule | Element | p | $e^-$ | n |
|---|---|---|---|---|
| $N_2$ | Nitrogen (1 × 2) | 2 × 7 = 14 | 2 × 7 = 14 | 2 × 7 = 14 |
| Total | | **14** | **14** | **14** |

**(d) In case of Charge on Molecule**

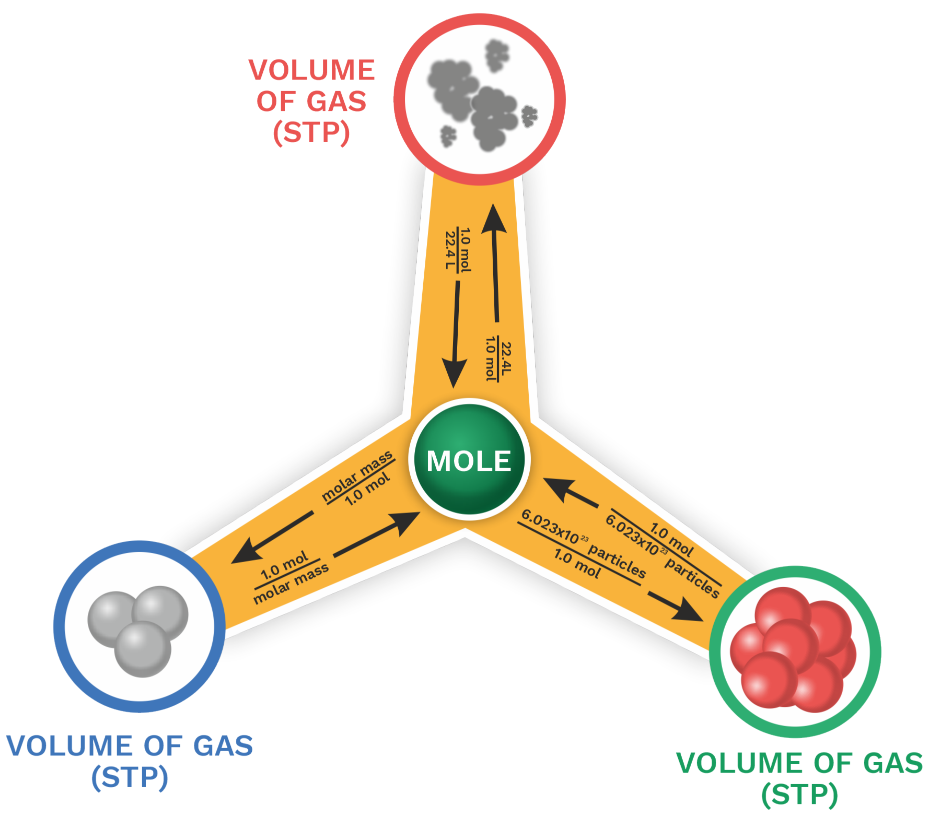
**Ex:** $(NH_4)^+$

Total number of atoms in $(NH_4)^+ = 5$

(No. of N-atoms = 1, No. of H-atoms = 4)

| Molecule | Element | p | $e^-$ | n |
|---|---|---|---|---|
| $(NH_4)^+$ | Nitrogen | 7 | 6 | 7 |
| | Hydrogen (1 × 4) | 4 | 4 | 0 |
| Total | | **11** | **10** | **7** |

## CALCULATION OF MOLES



- Amount of substance which consist of Avogadro's number ($6.022 \times 10^{23}$) of atoms if the substance is atomic or Avogadro's number ($6.022 \times 10^{23}$) of molecules.

  OR

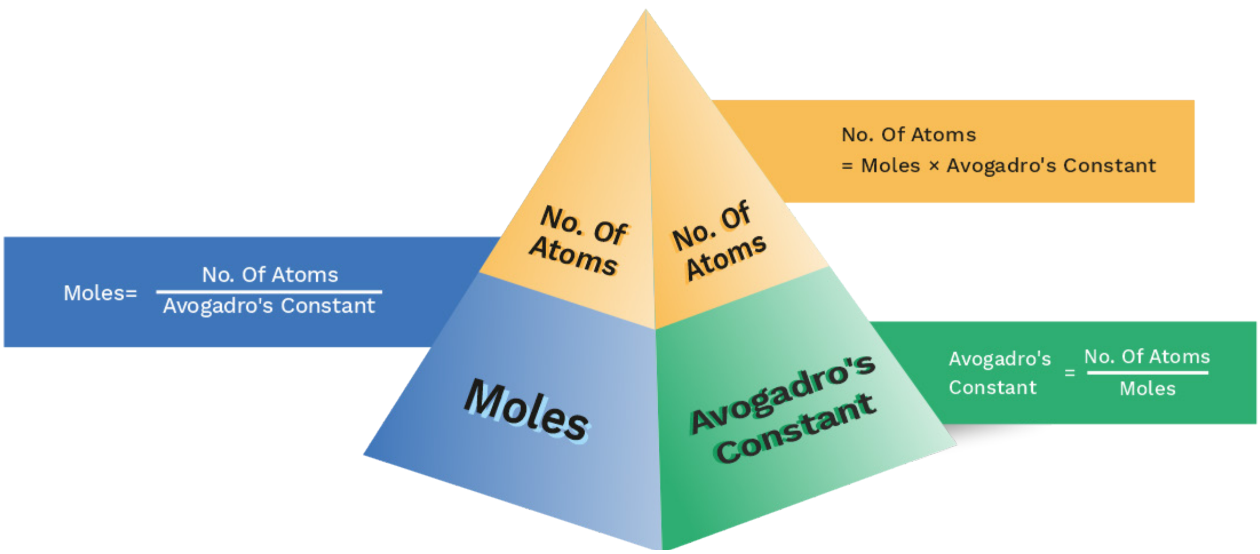
  In case of gaseous substance mole is the amount of gas which has a volume of 22.4 litres at STP.

**Previous Year's Question**

Which has maximum molecules?

**[AIPMT]**

(1) 7 g $N_2$ (2) 2 g $H_2$
(3) 16 g $NO_2$ (4) 16 g $O_2$




**Examples :**

**(a) In case of atoms :**

- 1 Mole of carbon atoms = $N_A$ carbon atoms

  OR

  1 Mole of carbon atoms = $6.022 \times 10^{23}$ carbon atoms
- 1 Mole of nitrogen atoms = $N_A$ nitrogen atoms = $6.022 \times 10^{23}$ Nitrogen atoms

**(b) In case of ions :**

- 1 Mole of $O^{2-}$ = $N_A$ $O^{2-}$ ions = $6.022 \times 10^{23}$ $O^{2-}$ ions
- 1 Mole of $N^{3-}$ = $N_A$ $N^{3-}$ ions = $6.022 \times 10^{23}$ $N^{3-}$ ions
- 1 Mole of $NH_4^+$ = $N_A$ $NH_4^+$ ions = $6.022 \times 10^{23}$ $NH_4^+$ ions

**Previous Year's Question**

Suppose the elements X and Y combine to form two compounds $XY_2$ and $X_3Y_2$. When 0.1 mole of $XY_2$ weighs 10 g and 0.05 mole of $X_3Y_2$ weighs 9g, the atomci weights of X and Y are

**[NEET-2016]**

(1) 40, 30 (2) 60, 40
(3) 20, 30 (4) 30, 20

**(c) In case of molecules :**

- 1 Mole of $N_2$ = $N_A$ $N_2$ molecules = $6.022 \times 10^{23}$ $N_2$ molecules
- 1 Mole of $O_2$ = $N_A$ $O_2^-$ molecules = $6.022 \times 10^{23}$ $O_2$ molecules
- 1 Mole of $CO_2$ = $N_A$ $CO_2$ molecules = $6.022 \times 10^{23}$ $CO_2$ molecules

**(d) In case of Acid/Base/Salt/Double salts :**

- 1 Mole of HCl = $N_A$ HCl = $6.022 \times 10^{23}$ HCl
- 1 Mole of NaOH = $N_A$ NaOH = $6.022 \times 10^{23}$ NaOH
- 1 Mole of $NH_4OH$ = $N_A$ $NH_4OH$ = $6.022 \times 10^{23}$ $NH_4OH$
- 1 Mole of NaCl = $N_A$ NaCl = $6.022 \times 10^{23}$ NaCl
- 1 Mole of $FeSO_4.(NH_4)_2SO_4.6H_2O$ = $6.022 \times 10^{23}$ $FeSO_4.(NH_4)_2SO_4.6H_2O$

**Previous Year's Question**

The number of moles of oxygen in one litre of air containing 21% oxygen by volume, under standard conditions, is

**[AIPMT]**

(1) 0.0093 mol
(2) 2.10 mol
(3) 0.186 mol
(4) 0.21 mol

**Q8** How many g atoms are there in on atom?

**A8** $6.023 \times 10^{23}$ atoms = 1 g atom = 1 mole

$$1 \text{ atom} = \frac{1}{6.023 \times 10^{23}}$$
$$= 1.66 \times 10^{-24} \text{ g atom or mol}$$

**Q9** From 200 mg of $CO_2$, $10^{21}$ molecules are removed. How many grams and moles of $CO_2$ are left.

**A9** 44 g of $CO_2$ = 1 mol = $6.023 \times 10^{23}$ molecules

$\therefore$ $6.023 \times 10^{23}$ molecules = 44 g of $CO_2$

$$10^{21} \text{ molecules} = \frac{40 \times 10^{21} \times 10^{3}}{6.023 \times 10^{23}}$$

Weight of $CO_2$ left = 200 – 73.05 = 126.9 mg $= \dfrac{126.9}{10^3} = 0.1269$ g

44 g of $CO_2$ = 1 mol

0.1269 g of $CO_2$ = $\dfrac{1}{44} \times 0.1269 = 0.0028$ mol

**Q10** How many moles of O are present in 4.9 g of $H_3PO_4$?

**A10** Molecular weight of $H_3PO_4$ = 1 × 3 + 31 + 16 × 4 = 98 g

98 g = 1 mole of $H_3PO_4$ = 4 mole of O

4.9 g of $H_3PO_4$ $= \frac{4}{98} \times 4.9$ = 0.2 mole of O

**Q11** Calculate the number of moles in each of the following:

(i) 11g of $CO_2$

(ii) $3.01 \times 10^{22}$ molecules of $CO_2$

(iii) 1.12 L of $CO_2$ at STP

**A11** (i) 44g of $CO_2$ = 1 mol

11g of $CO_2$ $= \frac{1}{44} \times 11 = \frac{1}{4} = 0.25$ mol

(ii) $6.023 \times 10^{23}$ molecuels = 1 mol

$3.01 \times 10^{22}$ molecules $= \frac{1 \times 3.01 \times 10^{22}}{6.023 \times 10^{23}} = 0.05$ mol

(iii) 22.4L of $CO_2$ = 1 mol

1.12L of $CO_2$ $= \frac{1 \times 1.12}{22.4} = 0.05$ mol

**Q12** What is the volume occupied by one $CCl_4$ molecule at 20°C? Density of $CCl_4$ is 1.6 g $mL^{-1}$ at 20°C.

**A12** $CCl_4$ molecule is liquid.

Mw of $CCl_4$ = 12 + 4 × 35.5 = 154 g

1 molecule of $CCl_4$ $= \frac{154}{6.023 \times 10^{23}} = 25.56 \times 10^{-23}$ g

$$\text{Volume of one molecule} = \frac{\text{mass}}{\text{density}}$$

$$= \frac{25.56 \times 10^{-23}}{1.6}$$

$$= 1.598 \times 10^{-22} \text{ mL or cm}^3$$

**Q13** The volume of a drop of water is 0.04 mL. How many $H_2O$ molecules are there in a drop of water? d = 1.0 g $mL^{-1}$.

**A13** Volume of 1 drop of $H_2O$ = 0.04 mL

Weight of $H_2O$ = Volume × Density = 0.04 × 1 = 0.04 g

1 mole of $H_2O$ = 18 g = $6.023 \times 10^{23}$ molecules

$$\therefore 0.04\ \text{g} = \frac{6.023 \times 10^{23} \times 0.04}{18} = 1.3384 \times 10^{21}\ \text{molecules}$$

## PRECENTAGE COMPOSITION

- The percentage fo 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 c

It is calculated as follows:

- First calculated the molecular mass of the compound from its formula by adding the atomic masses of the elements present in 100 parts by mass of the compound.
- Then calculate the percentage of the element or constituents by using the relation:

**Previous Year's Question**

An organic Compound contains carbon, hydrogne and oxygen. its elemental analysis gave C, 38.71% and H, 9.67%. The empirical formula of the compound would be

**[AIPMT-2008]**

(1) CHO (2) $CH_4O$

(3) $CH_3O$ (4) $CH_2O$

$$\text{Mass\% of an element} = \frac{\text{Mass of element × 100 in the compound}}{\text{Molar mass of the compound}}$$

**Q14** Calculate the percentage composition of various elements in the following compoud : Blue vitriol ($CuSO_4.5H_2O$).

**A14** Molar mass of $CuSO_4.5H_2O$ = 63.5 + 32 + 4 × 16 + 5 × 18 = 249.5 g

$$\text{Mass\% of Cu} = \frac{63.5 \times 100}{249.5} = 25.45\%$$

$$\text{Mass\% of S} = \frac{32 \times 100}{249.5} = 12.82\ \%$$

$$\text{Mass\% of O} = \frac{16 \times 9 \times 100}{249.5} = 57.71\ \%$$

$$\text{Mass\% of H} = \frac{10 \times 1.008 \times 100}{249.5} = 4.040\ \%$$

## EMPIRICAL AND MOLECULAR FORMULAE

### Empirical Formula

- Empirical formula gives the simplest whole number ratio of the various atoms present in a compound.
- For example, the empirical formula of glucose is $CH_2O$, that of benzene is CH, and that of hydrogen peroxide is OH.

### Molecular Formula

- It represents the exact number of different types of atoms present in a molecule of a compound.
- For example, molecular formual of glucose is $C_6H_{12}O_6$, that of hydrogen peroxide is $H_2O_2$.

### Relation Between Empirical and Molecular Formulae

- The molecualr formula of a compound is a simple whole number multiple of its empirical formula,

  **Molecualr formual = n × Empirical formula**

  where n is any integer, e.g., 1, 2, 3, ..., etc.

  The value of n is obtained as

$$n = \frac{\text{Molar mass}}{\text{Empirical formula mass}}$$

### Molar Mass of a Volatile Compound

- It is determined by Victor Meyer's method. It is based on the principle that 22.4 L of vapours of a volatile compound at STP have mass equal to the gram molecualr mass or by the relation given below.

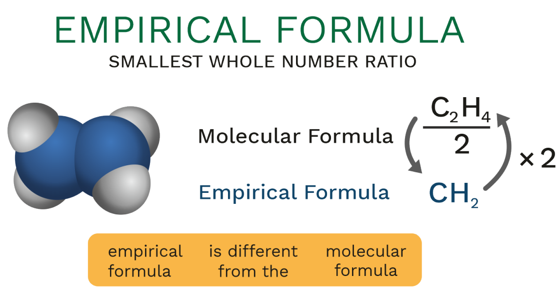
  **Molar mass = 2 × Vapour density**

### Calculation of the Empirical and Molecular Formulae

- **Conversion of mass percent to grams.**
- **Convert into number of moles of each element.**
- **To calculate the simplest whole number.**
- **Writing empirical formula.**
- **Writing molecular formula.**

**Rack your Brain** 

Can we apply percentage composition in case of non-stochiometric comopounds like $Fe_{0.95}O$?




**Concept Ladder** 

There can be certain substances where stoichiometry is not followed.

Example – $Fe_{0.95}O$.

**Q15** A substance, on analysis, gave the following percentage composition: Na = 43.4%, C = 11.3% and O = 45.3%. Calculate the empirical formula.
(Na = 23, C = 12, O = 16)

**A15**

| Element | Percentage composition | Atomic ratio | Least ratio |
|---|---|---|---|
| Sodium | 43.4 | $\frac{43.4}{23} = 1.89$ | $\frac{1.89}{0.94} = 2$ |
| Carbon | 11.3 | $\frac{11.3}{12} = 0.94$ | $\frac{0.94}{0.94} = 1$ |
| Oxygen | 45.3 | $\frac{45.3}{16} = 2.83$ | $\frac{2.84}{0.94} = 3$ |

Hence, the empirical formula is $Na_2CO_3$.

**Q16** Assuming the atomic weightof a metal M to be 56, find the empirical formula of its oxide containing 70.00% of M.

**A16**

| Element | Percentage composition | Atomic ratio | Least ratio | Whole number ratio |
|---|---|---|---|---|
| Metal | 70.00 | $\frac{70}{56} = 1.25$ | $\frac{1.25}{1.25} = 1$ | 2 |
| Oxygen | 30.00 | $\frac{30}{16} = 1.875$ | $\frac{1.875}{1.25} = 1.5$ | 3 |

Hence, the empirical formula is $M_2O_3$.

## STOICHIOMETRY AND STOICHIOMETRIC CALCULTIONS

- The word **stoichiometry** is taken from two Greek words — **stoicheion** (which means element) and **metron** (meaning measurement).
- It deals with the calculation of moles, molecules masses and sometimes volumes of the reactants and products involved in a **balanced chemical equation**.
- The coefficients of the balanced chemical equations are called **stoichiometric coefficients**.
- Stoichiometric coefficients represent the number of moles and molecules of reactants and products in a balanced chemical equations.

**Concept Ladder**

Stoichiometry uses all the laws of chemical combination during calculations.

**Q17** Find out the volume of $CO_2$ produced by the thermal decomposition of 5 mole of calcium carbonate at STP?

(1) 22.4 litre (2) 2 × 22.4 litre
(3) 5 × 22.4 litre (4) 3 × 22.4 litre

**A17** (3)

$$CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$$

1 mole 1 mole 1 mole

1 mol $CaCO_3$ : 1 mol $CO_2$

5 mol $CaCO_3$ : 5 mol $CO_2$

At STP 5 mole $CO_2$ will have a volume of 5 × 22.4 litre

**Q18** Find out the mass of CaO, obtained from the thermal decomposition of 5 gm of $CaCO_3$?

(1) 5.6 gm (2) 2.8 gm (3) 3.6 gm (4) 8.6 gm

**A18** (2)

$$CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$$

1 mol $CaCO_3$ : 1 mol CaO

100 g $CaCO_3$ : 56 g CaO

$$1\ g\ CaCO_3 : \frac{56}{100}\ g\ CaO$$

$$5\ g\ CaCO_3 : \frac{56}{100} \times 5\ g\ CaO$$

$$= 2.8\ gm$$

**Q19** Find out the volume of $CO_2$ obtained at NTP by the thermal decomposition 5gm of $CaCO_3$?

(1) 22.4 litre (2) 2.245 litre
(3) 33.6 litre (4) 1.12 litre

**A19** (4)

$CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$

$$n_{CaCO_3} = \frac{5}{100}$$

$$n_{CO_2} = n_{CaCO_3} = \frac{5}{100} = \frac{1}{20}$$

$$V_{CO_2} = \frac{1}{20} \times V_m = \frac{22.4}{20} = 1.12 \text{ Ltr}$$

**Q20** Find out the volume of $O_2$ obtained from the thermal decomposition of 0.1 mole of Potassium chlorate ($KClO_3$) at NTP ?

(1) 1.12 Litre (2) 2.24 Litre
(3) 3.36 Litre (4) 4.48 Litre

**A20** (3)

$2KClO_3 \longrightarrow 2KCl + 3O_2$

2 moles of $KClO_3$ produces 3 moles of $O_2$ according to the above chemical reaction.
So, 1 mol $KClO_3$ forms 3/2 mol $O_2$.
$\therefore$ 0.1 mole $KClO_3$ will give = 0.3/2 × 22.4 litre of $O_2$ = 3.36 litre

**Q21** Find out the mass of $O_2$ obtained from 90 kg of water?

(1) 70 kg (2) 80 kg
(3) 100 kg (4) 50 kg

**A21** (2)

$2H_2O \longrightarrow 2H_2 + O_2$
2 mole $H_2O$ forms 1 mole of $O_2$
2 × 18 g of $H_2O$ forms 32 g of $O_2$ from above reaction.

1 kg of $H_2O$ will form $\frac{32}{2 \times 18}$ kg of $O_2$

So, from 90 kg of $H_2O$ = $\frac{32 \times 90}{2 \times 18}$ kg of $O_2$ = 80 kg of $O_2$

**Q22** Calculate the weight of CaO obtained by heating of 200 kg of 95 % pure lime stone

(1) 212.8 kg (2) 106.4 kg

(3) 318.4 kg (4) 418.4 kg

**A22** (2)

Lime stone sample 95 % purity 200 kg

of sample = Pure $CaCO_3$ = $200 \times \frac{95}{100} = 190\,kg$

$CaCO_3\,(s) \longrightarrow CaO\,(s) + CO_2\,(g)$

1 mole of $CaCO_3$ produces 1 mole of CaO

100 gm of $CaCO_3$ will produces 56 gm of CaO

So, from 190 kg $CaCO_3$ = $\frac{56}{100} \times 190\ kg\ \ CaO = 106.4\ kg$

**Q23** Find out the weight of iron which will be converted into its oxide ($Fe_3O_4$) by the action of 18 g of steam?

(1) 21 gm (2) 42 gm

(3) 63 gm (4) 84 gm

**A23** (2)

From the compound given the following can be deduced

| | | | |
|---|---|---|---|
| Molar ratio | 3 mol Fe | : | 4 mol $H_2O$ |
| Mass ratio | 3 × 56 g Fe | : | 4 × 18 g $H_2O$ |

Using stoichiometry $m_{Fe} = \frac{3 \times 56 \times 18}{4 \times 18}$ = 42 gram

**Q24** Calculate the amount of water (g) produced by the combustion of 16g of methane.

**A24** $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$

(i) 16 g of Methane is equal to one mole.

(ii) From above equation, a mol of $CH_4(g)$ gives 2 mole of $H_2O(g)$.

So, mass of $H_2O$ = 2 × 18 = 36g

**Ideal gas equation :**

- $PV = nRT$
  $P$ = Pressure of the gas
  $V$ = Volume of the gas
  $n$ = mol of gas
  $R$ = universal gas constant
  $T$ = temperature in K
- $V_{gas} = V_{vesel}$
  $V_{gas}$ = Free available space for motion
  $R$ = universal gas constant $\quad R = \dfrac{PV}{nT}$

$$R = \frac{0.0821 \text{ atm L}}{\text{mol K}} \quad \text{or} \quad R = \frac{1.99 \text{ Cal}}{\text{mol K}}$$

$$\text{or} \quad R = \frac{8.314 \text{ Joule}}{\text{mol K}}$$

- $PV = nRT$ at constant P, T

$$\uparrow \boxed{V \propto n} \uparrow \qquad \boxed{\frac{V_1}{V_2} = \frac{n_1}{n_2}}$$

- $\boxed{\text{STP / NTP}}$ {Standard temp. & Pressure}
- $\boxed{P = 1 \text{ atm} \quad T = 273 \text{ K}}$ {Normal temp. & Pressure}

## Important points

**Conversion Factor for gaseous properties :**

**Pressure**

1 atm = 760 mm
= 760 torr
= 1.01325 bar = 101325 Pa

1 bar = $10^5$ Pa

**Volume**

1 $m^3$ = 1000 L
1 $dm^3$ = 1 L
1 $cm^3$ = 1 cc = 1 ml

**Temperature**

$T_{Kelvin} = 273 + T_{oC}$

$T_{Fahrenheit} = \frac{9}{5} T_{oC} + 32$

1 L atom = 101.325 J

**Q25** Find out the molar volume of an ideal gas at STP/ NTP?

**A25** $PV = nRT \qquad V = \dfrac{nRT}{P}$

At STP: P = 1 Bar, T = 273.15 K

For n = 1 mole and the value of $R = \dfrac{0.0821 \text{ atm L}}{\text{mol K}}$

Using ideal gas equation

$$V = \frac{nRT}{P} = \frac{1 \times 0.0821 \times 273.51}{0.987} = 22.7 \text{ L}$$

Volume of 1 mol of an ideal gas at 1 atm = 22.4 L

### Limiting Reagent

- If the reactants are not taken in the stoichiometric ratios then the reactant which is less than the required amount determines how much product will be formed.

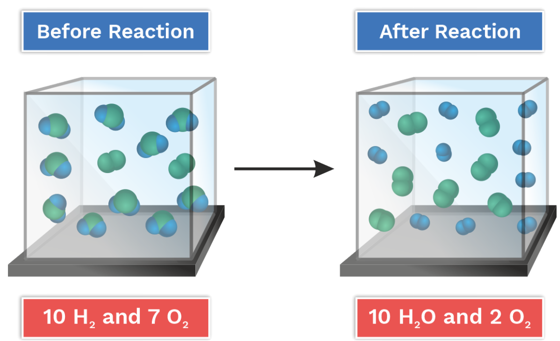
### Excess Reagent

- The substance which does not get consumed completely is known as excess reagent.
- The reactant present in excess is called the Excess Reagent.

**Ex:** If we burn carbon in air (which has an infinite supply of oxygen) then the amount of $CO_2$ being produced will be governed by the amount of carbon taken. In this case, Carbon is the LR and $O_2$ is the ER.

**Definition**

During a chemical reaction a substance gets consumed completely. This substance is known as limiting reagent.




**Q26** The number of litres of air required to burn 8 litres of $C_2H_2$ is approximately?

| | |
|---|---|
| (1) 40 | (2) 60 |
| (3) 80 | (4) 100 |

**A26** For a chemical reaction of the type given below :

Given mole at t = 0

$$\begin{array}{ccccccc} 1\,A & + & 2\,B & \longrightarrow & C & + & 2\,D \\ 5\text{ mole} & & 12\text{ mole} & & 0 & & 0 \end{array}$$

At the completion of reaction the following can be seen, when 12 moles of B reacts with 5 moles of A.

$$\begin{array}{ccccccc} A & + & 2\,B & \longrightarrow & C & + & 2\,D \\ [5-5]\text{ mole} & & [12-10] & & 5\text{ mole} & & 10\text{ mole} \\ 0 & & 2\text{ mol excess} & & \text{max} & & \text{max} \end{array}$$

In this case, A gets consumed completely so it behaves as limiting reagent for the reaction

$$C_2H_2 + \frac{5}{2}O_2 \longrightarrow 2CO_2 + H_2O$$

$$1\text{ ml} : \frac{5}{2}\text{ml}$$

$$8\text{ ml} : \frac{8 \times 5}{2}\text{ml}$$

8 ml : 20 ml of $O_2$

$20 \times 5 \simeq V_{air}$

$V_{air} = 100$ ml

## PERCENT YIELD

- The amount of product formed by a chemical reaction is less than the amount predicted by theoretical calculations.
- The ratio of the amount of product formed to the amount predicted when multiplied by 100 gives the percentage yield.

**Rack your Brain**

Why experimental yield is always less than theoretical or calculated yield?

$$\text{Percentage Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100$$

**Q27** For the reaction, $CaO + 2HCl \longrightarrow CaCl_2 + H_2O$.
1.12 gram of CaO si reacted with excess of hydrochloric acid and 1.85 gm $CaCl_2$ is formed. What is the % yield of the reaction?

**A27** $CaO + 2HCl \longrightarrow CaCl_2 + H_2O$

56 gm CaO will produce 111 gm $CaCl_2$

1.12 gram of CaO will produce $\rightarrow \frac{111}{56} \times 1.12 = 2.22$ gm

Thus Theoretical yield = 2.22 gm

Actual yield = 1.85 gm

$$\%\ \text{yield} = \frac{1.85}{2.22} \times 100 = 83.33\%$$

## PERCENTAGE PURITY

- Depending upon the mass of product, the equivalent amount of reactant present is determined with the help of given chemical equation. When we known the actual amount of reactant taken and the amount calculated with the help of a chemical equation, the purity can be determined.

**Rack your Brain**

Percentage purity is so important in our daily life. How?

$$\text{Percentage purity} = \left[\frac{\text{Amount of reactant calculated form the chemical equation}}{\text{Actual amount of reactant taken}}\right] \times 100\%$$

**Q28** Calcualte the amount of (CaO) in kg that can be produced by heating 200 kg lime stone taht is 90% pure $caCO_3$.

**A28** Mass of Pure $CaCO_3 = \frac{200 \times 90}{100} = 180$ kg

$CaCO_3 \longrightarrow CaO + CO_2$

100 kg    56 kg

180     x

$\frac{100}{180} = \frac{56}{x} \Rightarrow x = 100.8$ kg

### Eudiometry

- It is the special part of stoichiometry.
- In this, we deal with absorption of gases during a chemical process.

**Definition**

The process of determining the constituents of a gaseous mixture by means of the eudiometer, or for ascertainnig the purity of the air or the amount of oxygen in it.

### Some Absorbents of Gases

The absorbent which is used for specific gas is listed below:

| Absorbent | Gas or gases absorbed |
|---|---|
| Turpentine oil | $O_3$ |
| Alkaline pyrogallol | $O_2$ |
| Ferrous sulphate solution | NO |
| Heated magnesium | $N_2$ |
| Heated palladium | $H_2$ |
| Ammonical coprous chloride | $O_2$, CO, $C_2H_2$ |
| Copper sulphate solution | $H_2S$, $PH_3$, $AsH_3$ |
| Conc. $H_2SO_4$ | $H_2O$ i.e., moisture, $NH_3$ |
| NaOH or KOH solution | $CO_2$, $NO_2$, $SO_2$, $X_2$, all acidic oxides |

**Q29** 90 ml of pure dry $O_2$ is subjected to electric discharge, if only 10 % of it is converted into $O_3$, volume of the mixture of gases ($O_2$ & $O_3$) after the reaction will be ______and after passing through turpentine oil will be ________.

(1) 84 ml, 78 ml (2) 81 ml, 87 ml
(3) 78 ml, 84 ml (4) 87 ml, 81 ml

**A29**

| | $3O_2 \longrightarrow$ | $2O_3$ |
|---|---|---|
| Initially | 90 ml | 0 |
| After reaction | 90 – 9 ml | 6 ml |

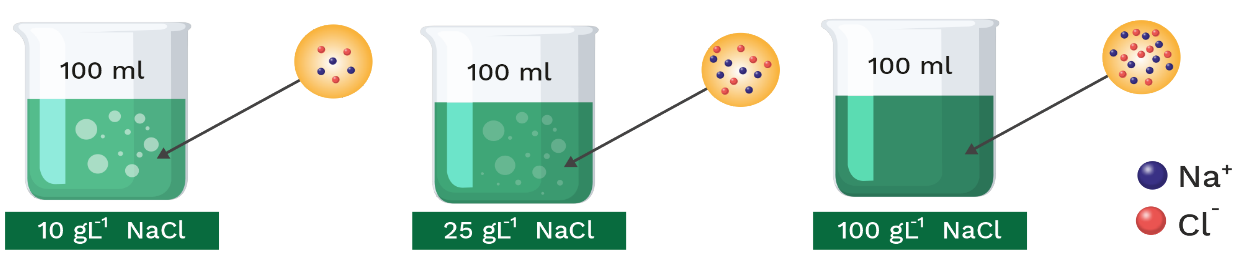
$V_{mix}$ = 81 ml $O_2$ + 6 ml $O_3$ = 87 ml
Turpentine oil (It absorbs $O_3$ gas)
$V_{remaining}$ = 81 ml of $O_2$

## CONCENTRATION TERMS

- A solution is defined as the homogeneous mixture of two or more substances, composition of which may vary within the limits.
- Solution is a special type of mixture in which substances are intermixed so intimately that they cannot be observed as separated components.
- The substance which is dissolved is called solute while medium in which solute is dissolved to get a homogeneous mixture is called the solvent.
- A solution is termed as binary and ternary if it consists of two and three components

**Rack your Brain**

How does strength of solution change by increase in temperature?



respectively.

**Percentage**

- It refers tot he amount of the solute per 100 parts of the solution. It can also be called as parts per hundred (pph). It can be expressed be any of following four methods:

**(i) Weight by weight percentage (%w/w)**

$$= \frac{\text{Wt. of solute(g)}}{\text{Wt. of solution(g)}} \times 100$$

e.g., 10% $Na_2CO_3$ solution w/w means 10g of $Na_2CO_3$ is dissolved in 100 g of the solution.

**(ii) Weight by volume percent (%w/v)**

$$= \frac{\text{Wt. of solute(g)}}{\text{Volume of solution(cm}^3)} \times 100$$

e.g., 10% $Na_2CO_3$ (w/v) means 10 g $Na_2CO_3$ is dissolved in 100 $cm_3$ of solution.

**(iii) Volume by bolume percent (%v/v)**

$$= \frac{\text{Volume of solute(cm}^3)}{\text{Volume of solution(cm}^3)} \times 100$$

e.g., 10% ethanol (v/v) means 10 $cm^3$ of ethanol dissolved in 100 $cm^3$ of solution.

**(iv) Volume by weight percent (%v/w)**

$$= \frac{\text{Volume of solute(cm}^3)}{\text{Wt. of solution(g)}} \times 100$$

e.g, 10% ethanol (v/w) means 10$cm^3$ of ethanol dissolved in 100g of solution.

**Concept Ladder**

$$M = \frac{\%\text{by mass} \times d_{(\text{solution})} \times 10}{M_{w_2}}$$

Where,
M = Molarity of solution
$d_{solution}$ = density of solution
$M_{w_2}$ = Molar mass of solute

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Is there any effect of change in temperrature on %v/v?

**Concept Ladder**

%v/v is used for solution which contains both solute and solvent in liquid state.

### Parts per million (ppm) and parts per billion (ppb)

- When a solute is present in very small quantity, it is convenient to express the concentration in parts per million and parts per billion.
- It is the number of parts of solute per million ($10^6$) or per billion ($10^9$) parts of solution.
- It is independent of the temperature.

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

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

**Rack your Brain**

What is the application of ppm, ppb or ppt in daily life?

**Q30** Calculate the parts per million of $SO_2$ gas in 250 mL water containing 5 × $10^{04}$ g of $SO_2$ gas.

**A30** Mass of $SO_2$ gas = 5 × $10^{-4}$g;

Mass of $H_2O$ = Volume × Density = 250 $cm^3$ × 1 g $cm^{-3}$ = 250 g

∴ Parts per million of $SO_2$ gas $= \dfrac{5 \times 10^{-4}}{250\text{ g}} \times 10^{-6} = 2$

### Molarity (M)

- Molarity is defined as the number of moles of the solute per liter of solution. Unit of molarity is mol/$dm^3$ or mol/liter.
- For example, one molar (1 M) solution of sugar means the solution contains 1 mole of sugar per litre of the solution. Solution in terms of molarity is generally expreesd as,
- Mathematically, molarity can be calculated by following formulas:

**Rack your Brain**

Can molarity be used to express concentration for ionic solution?

(i) $$M = \frac{\text{No. of moles of solute(n)}}{\text{Vol. of solution in litres}} = \frac{\text{Wt. of solute(gm)}}{\text{gm mol.wt. of solute}} \times \frac{1000}{\text{wt. of solution(ml)}}$$

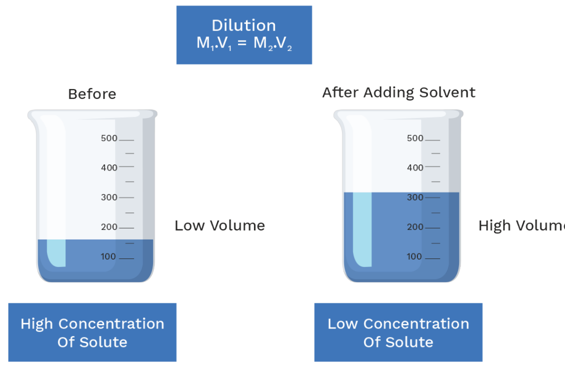
(ii) If molarity and volume of the solution are changed from $M_1$, $V_1$ to $M_2$, $V_2$.

Then, $M_1V_1 = M_2V_2$

(iii) (iii)In the balanced chemical equation, if n1 moles of reactant-1 react with n2 moles of reactant-2. Then,

$$n_1A + n_2B \rightarrow \text{Product}$$

$$\frac{M_1V_1}{n_1} = \frac{M_2V_2}{n_2}$$




(iv) (iv)If two solutions of same solute are mixed then molarity (M) of resulting solution

$$M = \frac{M_1V_1 + M_2V_2}{(V_1 + V_2)}$$

**Q31** A bottle of commercial sulphuric acid (density 1.787 g $ml^{-1}$) is labelled as 86% by weight. What is the molarity of acid?

**A31** Molarity of $H_2SO_4 = \dfrac{\text{Wt. of } H_2SO_4 \text{ in 1L solution}}{\text{mol. wt. of } H_2SO_4}$

But wt. of given $H_2SO_4$ per litre $= \dfrac{86}{100} \times 1.787 \times 1000 = 1536.82$ g

Hence molarity of $H_2SO_4 = \dfrac{1536.82}{98} = 15.68$ mol $L^{-1}$

**Q32** A sample contains $I_2$ and benzene. The mole fraction of $I_2$ = 0.2. Calculate molarity of solution if

(i) density of solution is d gm/ml

(ii) density of $I_2$ & benzene are $d_{I2}$ & $d_{benzene}$

**A32** (i) M = 1.77 d

(ii) $M = \dfrac{0.2}{\left(\dfrac{50.8}{d_{I_2}} + \dfrac{62.4}{d_{benzene}}\right)} \times 1000$

### Formality (F)

- Formality of solution may be defined as the number of gram formula unis of the ionic solute dissolved per litre of the solution.

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

$$= \frac{\text{Mass of ionic solute (g)}}{\text{gram formula unit mass of solute} \times \text{Volume of solution (l)}}$$

**Q33** What will be the formality of $KNO_3$ solution having strength equal to 2.02 g per litre?

**A33** Sterngth of $KNO_3$ = 2.02 $gL^{-1}$

$$\therefore \text{Formality of } KNO_3 = \frac{\text{strength in } gL^{-1}}{\text{g-formula wt. of } KNO_3} = \frac{2.02}{101} = 0.02 \text{ F}$$

### Molality (m)

- It is the number of moles of the solute per 1000 g of the solvent.
- Unit of molarity is mol/kg.
- Mathematically, molality can be calculated by the following formulae,

(i) $$m = \frac{\text{Number of moles of solute}}{\text{Weight of solvent in kg}} = \frac{\text{Number of moles of solute}}{\text{Weight of solvent in gm}} \times 1000$$

(ii) $$m = \frac{\text{Wt. of solute}}{\text{Mol. wt. of solute}} \times \frac{1000}{\text{Weight of solvent in gm}}$$

$$\text{Molality (m)} = \frac{M \times 1000}{1000d - MM_o}$$

**Q34** Calculate the Molarity and molality of a 98% by mass of $H_2SO_4$ solution having a density of 1.25 g/cc.

**A34** $H_2SO_4$ taken = 98%

100g of solution contains 98g $H_2SO_4$.

mass of solution = 100g

mass of solute, $H_2SO_4$ = 98g

mass of volvent = 100 – 98 = 2g = 0.002 kg

moles of solute, $H_2SO_4 = \frac{98}{98} = 1$

$$\text{Volume of solution} = \frac{\text{mass of solution}}{\text{density}} = \frac{100}{1.25} = 80\,\text{mL} = 0.08\,\text{L}$$

$$\text{Molarity, M} = \frac{\text{moles of solute}}{\text{volume of solution (L)}} = \frac{1}{0.08} = 12.5\ \text{M}$$

$$\text{Molarity, M} = \frac{\text{moles of solute}}{\text{mass of solution (kg)}} = \frac{1}{0.02} = 500\ \text{m}$$

## Mole fraction (X)

- It is defined as ratio of number of moles of a component to total moles of all components (solvent and solute) present in solution.
- It is denoted by the letter X.

Number of moles of component A is given by,

$$n_A = \frac{W_A}{M_A}$$

Number of moles of component B is given by,

$$n_B = \frac{W_B}{M_B}$$

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

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

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

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

$$X_A + X_B = \frac{n_A}{n_A + n_B} + \frac{n_B}{n_A + n_B} = 1$$

**Concept Ladder**

Relationship between molality, moles and molefraction.

$$m = \frac{n_2 \times 1000}{n_1 \times M_{w_1}} = \frac{x_2 \times 1000}{x_1 \times M_{w_1}}$$

$n_2$ = no. of moles of solute

$n_1$ = no. of moles of solvent

$M_{w_2}$ = molar mass of solvent

$x_2$ = molefraction of solute

$x_1$ = molefraction of solvent

**Q35** Find out the masses of acid and water requried to prepare 1 mle of $CH_3COOH$ solution of 0.3 mole fraction of $CH_3COOH$.

**A35** $X_{CH_3COOH} = 0.3$

$X_{H_2O} = 1 - 0.3 = 0.7$

$\text{Wt. of } CH_3COOH = X_{CH_3COOH} \times \text{mol. wt.}(CH_3COOH) = 0.3 \times 60 = 18 \text{ g}$

$\text{Wt. of water} = X_{H_2O} \times \text{mol. wt.}(H_2O) = 0.7 \times 18 = 12.6 \text{ g}$

**Q36** From 160 g of $SO_2$ (g) sample, $1.2046 \times 10^{24}$ molecules of $SO_2$ are removed then find out the volume of left over $SO_2$ (g) at STP.

**A36** Given moles = $\frac{160}{64} = 2.5$

Removed moles = $\frac{1.2046 \times 10^{24}}{6.023 \times 10^{23}} = 2$

Number of moles left = 0.5.

Remaining volume at STP = 0.5 × 22.4 = 11.2 lit..

**Q37** When x gram of a certain metal brunt in 1.5 g oxygen to give 3.0 g of its oxide. 1.20 g of the same metal heated in a steam gave 2.40 g of its oxide. shows the these result illustrate the law of constant or definite proportion

**A37** weight of metal = 3.0 – 1.5 = 1.5 g

So, weight of metal : weight of oxygen = 1.5 : 1.5 = 1 : 1

Similarly, in second case,

weight of oxygen = 2.4 – 1.2 = 1.2 g

so weight of metal : weight of oxygen = 1.2 : 1.2 = 1 : 1

so, them results illustrate law of constant proportion.

**Q38** A fresh $H_2O_2$ solution is labelled 11.2 V. This solution has the same concentration as a solution which is:

(1) 3.4% (w/w) (2) 3.4% (v/v)

(3) 3.4% (w/v) (4) None of these

**A38** V = 11.2

= M = 1 mol/L

Hence there are 34g $H_2O_2$ in 1 litre

i.e., 3.4g in 100mL = Hence 3.4% (wt./vol.)

**Q39** Density for 2 M CH3COOH solution is (1.2 g/mL)

Calculate

(1) Molality

(2) mole fraction of CH3COOH acid

(3) % (w/w) after the solution

(4) % (w/v) for the solution

**A39** Let volume & solution = 1 Litre

Molarity = $\frac{n_{solute}}{V_{solution}} = 2$

n solute = 2 ($CH_3COOH$) $\frac{\text{mass of solute}}{\text{molar mass of solute}} = 2$

mass of solute = 2 × 60 = 120 g.

mass of solution = volume of solution × density of solution = 1200 g

∴ mass of solvent = 1200 – 120 = 1080 g = 1.08 kg

moles of solvent = $\frac{1080}{18}$ = 60 mole

(1) Molality = 1.85 M

(2) $X_{CH_3COOH}$ = 0.0322

(3) % w/w = 10%

(4) % w/v = 12%

## Chapter Summary

- Atom is the fundamental unit of matter which is further indivisible i.e. atom can neither be created nor be destroyed.
- Actual mass of the mass of one atom or one molecule of a substance is called as actual mass.
- 1 amu = $\frac{1}{12}$ × mass of one C-12 atom = 1.66 × $10^{-24}$ g or 1.66 × $10^{-27}$ kg
- 1 mole of atoms is also termed as 1 gm-atom, 1 mole of ions is termed as 1 gm-ion and 1 mole of molecule termed as 1 gm - molecule
- Relation between the molecular formula and Empirical formula
- $n = \dfrac{\text{Molecular mass}}{\text{Empirical Formula mass}}$
- Vapour density = $\dfrac{\text{molecular mass}}{2}$
- **Limiting reagent :** Calculating amount of anyone product obtained taking each reactant one by one irrespective of other reactants. The one giving least product is limiting reagent.
- For reversible reaction, the actual amount of any limiting reagent consumed in such incomplete reaction is given by [% yield × given mole of limit reagent]
- Measuring volume is equivalent to counting number of molecules of gas

General Concentration term :

(a) Density = $\dfrac{\text{mass}}{\text{volume}}$, Unit : gm/cc

(b) Relative density = $\dfrac{\text{Density of any substance}}{\text{Density of reference substance}}$

(c) Specific gravity = $\dfrac{\text{Density of any substance}}{\text{Density of water at } 4°C}$

(d) Vapour density = $\dfrac{\text{Density of vapour at some temperature and pressure}}{\text{Density of } H_2 \text{ gas at same temperature and pressure}}$

- Concentration Terms

$$\% = \left(\frac{v}{V}\right) = \frac{\text{volume of solute}}{\text{volume solution}} \times 100$$

$$\text{Mole } \% = \frac{\text{Moles of solute}}{\text{Total moles}} \times 100$$

$$\text{Mole fraction}\left(X_a\right) = \frac{\text{moles of solute}}{\text{Total moles}}$$

$$\text{Molality}\left(M\right) = \frac{\text{Mole of soulte}}{\text{volume of solution in litre}}$$

$$\text{Molality}\left(m\right) = \frac{\text{moles of solute}}{\text{Mass of solvent(inkg)}}$$

$$\text{Parts per million}\left(ppm\right) = \frac{\text{mass of solute}}{\text{mass of solution}} \times 10^6$$