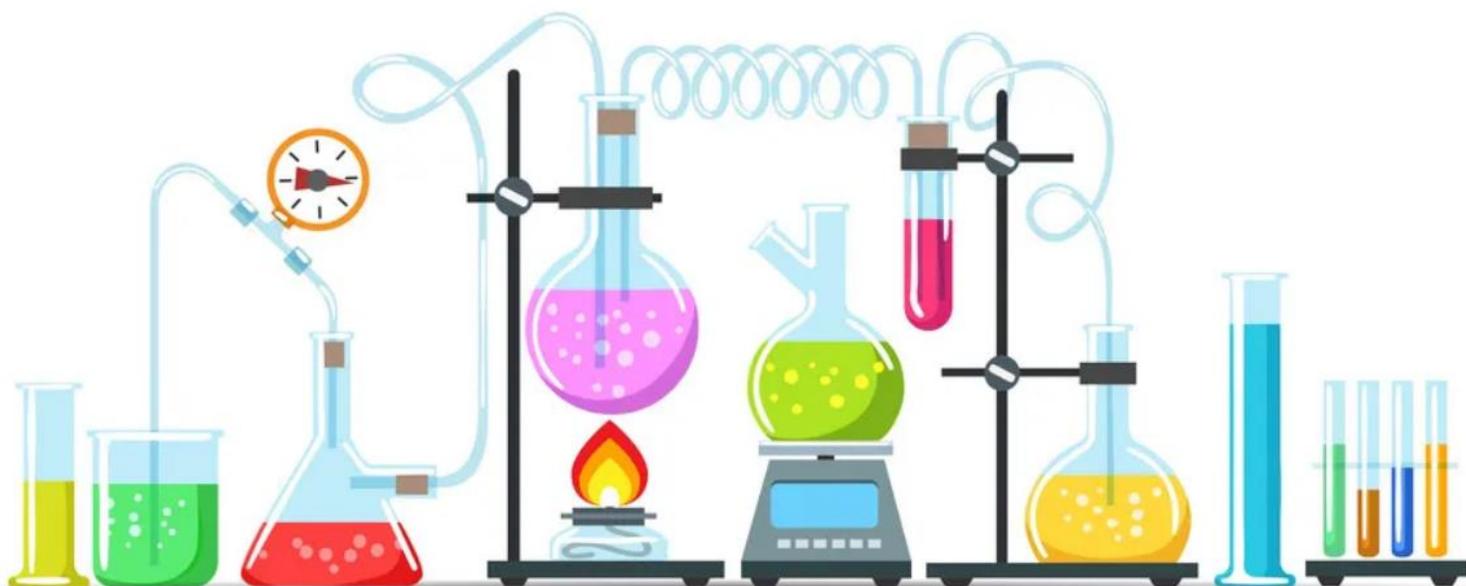# CHEMISTRY

CHAPTER 1: SOME BASIC CONCEPTS OF CHEMISTRY

SOME BASIC CONCEPTS OF CHEMISTRY

Introduction:

Chemistry is the science of molecules and their transformations which deals with the study of matter, its composition, the changes that matter undergoes and the relation between changes in composition and changes in energy. Chemistry plays an important role in meeting human needs for food, health care products.

Branches of Chemistry:

- **Organic Chemistry**- This branch deals with study of carbon compounds especially hydrocarbons and their derivatives.

- **Inorganic Chemistry**- This branch deals with the study of compounds of all other elements except carbon. It largely concerns itself with the study of minerals found in the Earth's crust.

- **Physical Chemistry**- The explanation of fundamental principles governing various chemical phenomena is the main concern of this branch. It is basically concerned with laws and theories of the different branches of chemistry.

- **Industrial Chemistry**- The chemistry involved in industrial processes is studied under this branch. Analytical Chemistry-This branch deals with the qualitative and quantitative analysis of various substances.

- **Biochemistry**- This branch deals with the chemical changes going on in the bodies of living organisms; plants and animals.

- **Nuclear Chemistry**- Nuclear reactions, such as nuclear fission, nuclear fusion, transmutation processes etc. are studied under this branch.
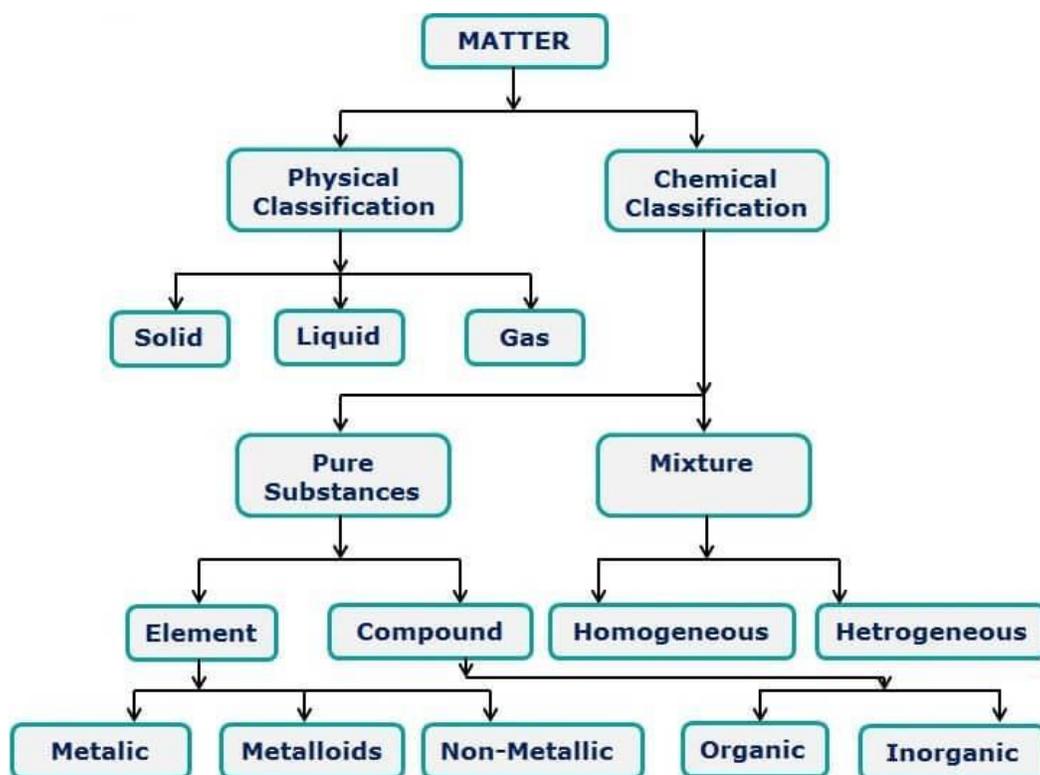
Importance of Chemistry:

Chemistry plays a central role in science and is often intertwined with other branches of science. Principles of chemistry are applicable in diverse areas, such as weather patterns, functioning of brain and operation of a computer, production in chemical industries, manufacturing fertilisers, alkalis, acids, salts, dyes, polymers, drugs, soaps, detergents, metals, alloys, etc., including new material.

Chemistry contributes in a big way to the national economy. It also plays an important role in

meeting human needs for food, healthcare products and other material aimed at improving the quality of life.

Matter

Matter is any thing that occupies space, has mass, offer resistance and can be perceived of directly by our senses. For example, book, pen, pencil, water, air, all living beings, etc.

```
                            MATTER

        Physical                      Chemical
        Classification                Classification

   Solid      Liquid      Gas

              Pure                    Mixture
              Substances

   Element    Compound         Homogeneous      Hetrogeneous

 Metalic   Metalloids   Non-Metallic      Organic    Inorganic
```

(1) Physical classification

(a) Solid

(b) Liquid

(c) Gases

Solid	Liquid	Gases
(1) They have fixed shape and volume	They do not have fixed shape but have fixed volume	They do not have fixed shape and volume

(2) They cannot be compressed	They cannot be compressed	They can be compressed easily
(3) They have high density	They have moderate density	They have low density
(4) They do not flow	They flow easily	They flow easily
(5) They do not fill their container	They do not fill their container	They fill their container
(6) The forces of attraction are strong	The forces of attraction are less strong than solids	The forces of attraction are weak.
(7) Kinetic energy is least	Kinetic energy is more than solids	Kinetic energy is maximum.
(8) Particles are closely packed	Particles are not close as in solids	Particles are much farther apart from one another.
(9) For Example : Chair, table, chalk, book	For Ex: Water, petrol, cold drinks	For Ex: Oxygen, nitrogen, helium

Vapours represent a gaseous state of a substance which is liquid at room temperature.

A substance which is in gaseous state at room temperature is called a gas.

For Ex: Ammonia is a gas but on heating water forms vapours.

(2) Chemical classification

All kinds of matter are classified into two types:

(a) Homogeneous

(b) Heterogeneous

Material is said to be **homogeneous** if it has uniform composition and identical properties throughout Or a material is said to be homogeneous if it consist of only one phase.

A material is said to be **heterogeneous** if it consists of a number of phases. The different phases are separated from each other by distinct boundaries.
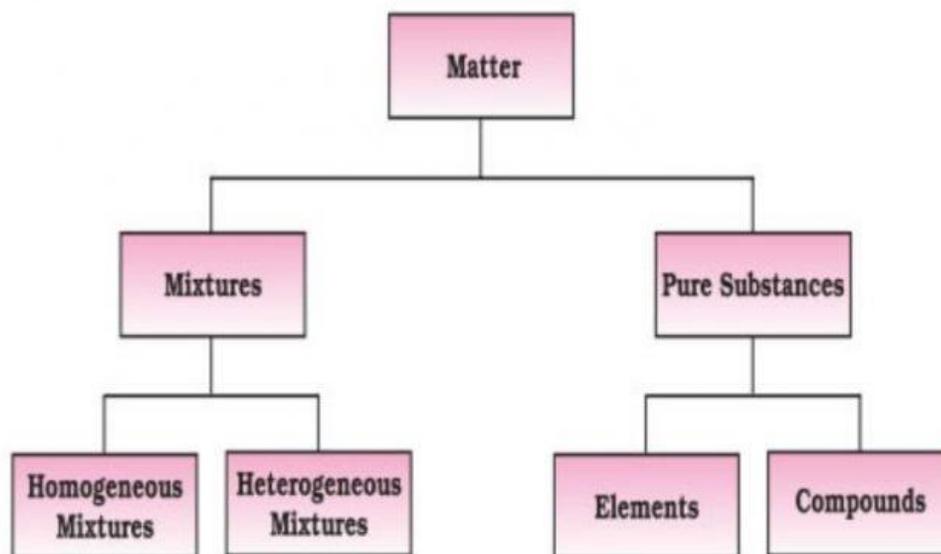
Element

Element is the purest form of matter. It is made up of only one type of atoms, ex.- carbon, iron, copper, oxygen etc.

Compound

Compound is the substance which is made up of two or more elements combined together in a fixed ratio by their weight e.g., carbon dioxide.

Mixture

Mixture is the substance which is made up of two or more substances in any ratio. e.g.,
Sugar + Water, Sodium Chloride + Water, Sand + Water

On the basis of composition, mixtures are of following type:

➤ **Homogeneous mixture**: The mixture which has uniform composition through out e.g., sugar solution.

➤ **Heterogeneous mixture**: The mixtures which do not have uniform composition through out. e.g. sand in water.

Atom

Atom is the smallest particle which may or may not exist free but takes part in chemical reaction. Atom word means not to be cut. Ex- H, Na, O etc.

International System of Units (S.I.)

The international system of units (in French Le Systeme International d' Unités - abbreviated as SI) was established in 1960 by the 11th general conference on weights and measures. SI system is a modification of metric system and has seven base units pertaining to the seven fundamental scientific quantities.

Base Physical Quantity	Symbol for Quantity	Name of SI Unit	Symbol for SI Unit
Length	l	metre	m
Mass	m	kilogram	kg
Time	t	second	s
Electric current	I	ampere	A
Thermodynamic temperature	T	kelvin	K
Amount of substance	n	mole	mol
Luminous intensity	I_v	candela	cd

Prefixes in SI system

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

Definition of SI Base Units

- **Metre**: The metre is the length of the path travelled by light in vacuum during a time interval of 1/(299792458) of a second.

- **Kilogram**: The kilogram is the unit of mass; it is equal to the mass of the international prototype of the kilogram.

- **Second**: The second is the duration of 9192631770 periods of the radiation corresponding to the transition between the two hyperfine levels of the ground state of the caesium-133 atom.

- **Ampere**: The ampere is that constant current which, if maintained in two straight parallel conductors of infinite length, of negligible circular cross-section, and placed 1 metre apart in vacuum, would produce between these conductors a force equal to 2×10^{-7} newton per metre of length.

- **Kelvin**: The kelvin, unit of thermodynamic temperature, is the fraction 1/273.16 of the thermodynamic temperature of the triple point of water.

- **Mole**: The mole is the amount of substance of a system which contains as many elementary entities as there are atoms in 0.012 kilogram of carbon-12; its symbol is "mol".

- **Candela**: The candela is the luminous intensity, in a given direction, of a source that emits monochromatic radiation of frequency 540×10^{12} hertz and that has a radiant intensity in that direction of 1/683 watt per steradian.

Uncertainty in Measurement

All scientific measurements involve certain degree of error or uncertainty. Scientific notations, significant figures and dimensional analysis help us in many ways in presenting of data and theoretical calculations.

Scientific Notation

It is an exponential notation in which any number can be represented in the form $N \times 10^n$ where n is an exponent having positive or negative values and N can vary between 1 to 10. Thus, 232.508 can be written as 2.32508×10^2 in scientific notation.

Precision and Accuracy

Precision refers to the closeness of various measurements for the same quantity. However, accuracy is the agreement of a particular value to the true value of the result. Let the true value of a quantity is 3.9 and its measurements taken by two boys are 3.6 and 3.8. Here 3.8 is more accurate as it is closer to the true value. Similarly 3.85 is more precise than 3.9.

Significant Figures

i. The total number of digits in measuring of any physical quantity with certainty is called significant figures. There are certain rules for determining the number of significant figures.

ii. All digits are significant except zero in the beginning of a number. For example, in 285 cm, there are three significant figures.

iii. Zeros to the left of the first non-zero digit are not significant if such zeros follow the decimal point. For example, 0.03 has one significant figure.

iv. Zeros to the right of the decimal point are significant. For example, 0.200 g has three significant figures.

v. Zeros between two non-zero digits are significant. Thus, 2.005 has four significant figures.

vi. Counting the numbers of object, for example, 2 balls or 20 eggs, have infinite significant figures as these are exact numbers and can be represented by writing infinite number of zeroes after placing a decimal i.e., 2 = 2.000000 or 20 = 20.000000.

Notes: In additions or subtractions, the final result should be reported to the same number of decimal places as that of the term with the least number of decimal places.

Laws of Chemical Combination

All chemical reactions take place according to certain laws. These laws are known as laws of **chemical combination**.

i. **Law of conservation of mass**: This law was put forth by Antoine Lavoisier in 1789. According to this, "It states that the total mass of reactants is equal to the total mass of the products".

ii. **Law of constant composition:** This law was given by, a French chemist, Joseph Proust. This law states that a chemical compound is always found to be made of same elements combined together in fixed proportion by weight. CO_2 can be prepared by number of methods but always 12g carbon react with 32g of oxygen.

iii. **Law of multiple proportions**: This law was proposed by Dalton in 1803. According to this law, When two elements combine to form two or more chemical compounds, then weight of one of the element which combines with a fixed weight of the other, bears a simple whole number ratio to one another. This is called the law of multiple proportions.
For example, The ratio of masses of oxygen in CO and CO_2 for fixed mass of carbon (12) is 16 : 32 = 1 : 2.

iv. **Law of reciprocal proportions**: It states that the ratio of weights of two elements A and B, which combine separately with the fixed weight of a third element C is either same or some simple whole number of the ratio of weights in which A and B combine directly with each other.
For example, ratio of masses of carbon and sulphur which combine with the fixed mass (32 parts) of oxygen is 12 : 32 or 3 : 8.

v. **Gay Lussac's Law of Gaseous Volumes**: This law was given by **Gay Lussac** in 1808. He observed that when gases combine or are produced in a chemical reaction they do so in a simple ratio by volume, provided all gases are at the same temperature and pressure.
For example, One volume of hydrogen and one volume of chlorine always combine to form two volumes of HCl gas.

vi. **Avogadro's Law**: In 1811, Avogadro proposed that equal volumes of all gases at the same temperature and pressure should contain equal number of molecules.

Dalton's Atomic Theory

John Dalton in 1808 published "A New System of Chemical Philosophy" in which he proposed atomic theory of matter. The main points of Dalton's atomic theory are as follows:

i. Matter is made up of extremely small, indivisible particles called atoms.

ii. Atoms of a given element are identical in all respect, i.e., they possess same size, shape, mass, chemical properties etc.

iii. Atoms of different elements are different in all respects, i.e., they possess different sizes, shapes, masses, chemical properties etc.

iv. Atoms of different elements may combine with each other in a fixed, simple, whole number ratio to form compounds.

v. Atoms can neither be created nor destroyed in a chemical reaction. Dalton's theory could explain the laws of chemical combination.

Atomic Mass and Molecular Mass

a) Atomic Mass: Atomic mass can be defined as a mass of a single atom which is measured in **atomic mass unit** (amu) or **unified mass** (u) where,

1 a.m.u. $= \dfrac{1}{12}$ **th** of the mass of one C-12 atom

b) Molecular Mass: Molecular mass is the sum of atomic masses of the elements present in a molecule. It is obtained by multiplying the atomic mass of each element by the number of its atoms and adding them together. **Molecular mass** expressed in grams is known as gram molecular mass.

Molecular mass of methane,

$(CH_4) = (12.011 \text{ u}) + 4 (1.008 \text{ u}) = 16.043 \text{ u}$

The Mole

One mole is the amount of substance that contains as many as entities as number of atoms in exactly 12.00g of C-12.
Number of carbon atoms in 12 g of C-12 $= 6.022 \times 10^{23}$

Chemical Formulae

Symbolic representation of compound is called chemical formula. It is of following types:

a) **Empirical Formula**: An **empirical formula** represents the simplest whole number ratio of various atoms present in a compound.

b) **Molecular Formula**: The molecular formula shows the exact number of different types of atoms present in a molecule of a compound.

Relationship between Empirical and Molecular Formula

Molecular formula = (Empirical formula) × n

Measurement of Concentration

The concentration of a solution reflects the relative proportion of solute and solvent present in the solution. The various concentration terms are,

1. **Weight percent** (% w/W) = (Weight of solute / Weight of solution) × 100.

2. **Volume percent** (% V/V) = (Volume of solute / Volume of solution) × 100.

3. **Molality (m)** – It is defined as number of moles of solute present in 1kg of solvent.

   m = {Number of moles of solute/Mass of solvent (in kg)} × 100.

4. **Molarity (M)** – It is defined as number of moles of solute present in 1 L of solution.

   M = {Number of moles of solute/Volume of solution (in litre)} × 100

5. **Mole fraction** : Suppose, n is the moles of solute and N is the moles of solvent, then,

   (i). Mole fraction of solute $(X_{solute}) = \dfrac{n}{n+N}$

   (ii). Mole fraction of solvent $(X_{solvent}) = \dfrac{N}{n+N}$

   $X_{solute} + X_{solvent} = 1$

6. **Normality**: It is defined as gram equivalent of solute dissolved in one litre solution.

$$N = \frac{\text{Gram equivalent of solute}}{\text{Volume of solution (litre)} \times 100}$$

Limiting Reagent

Limiting reagent is the reactant which is completely consumed in a reaction. To estimate the amount of product, limiting reagent should be known.

N_2(1 mole) + 3H_2(3 mole) $\longrightarrow$ 2NH_3(2 mole)

It means 1 mole of N_2 react with 3 mole of H_2 to produce 2 mole of NH_3.

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Physical Quantities

Physical Properties

Physical Properties are those which can be measured or observed without changing the identity or composition of the substance.

For example : Mass, volume, melting point, boiling point.

Chemical properties

Chemical properties are those in which a chemical change in the substance occurs.

The measurement of any physical quantity consists of two parts:

(a) The number

(b) The unit

For example: If an object weighs 4.5 kg it involves two parts:

4.5 is the number and kg is the unit.

A unit is defined as the standard of reference chosen to measure any physical quantity. French Academy of science in 1791 introduced a new system of measurement called metric system in which the different units of a physical quantity are related to each others as multiples of powers of 10.

For example: The improve System of units has been accepted internationally and is called International.

System of units or in short SI units (Systeme Internationale in French).

Seven base units

The seven basic physical quantities on which the international System of units is based, their symbols, the name of their units and the symbols of these units are given as:

Seven basic units

Measurement of Temperature

There are 3 scales of temperature:

(1) Degree Celsius

(2) Degree Fahrenheit

(3) Kelvin

The SI unit of Temperature is Kelvin

Relation between Degree kelvin and degree Celsius

Fahrenheit and Celsius relation

Relation between degree celsius and Degree Fahrenheit

Fahrenheit and Celsius relation

Measurement of Volume

The SI unit of volume is m^3

1 L =1000 mL =1000 cm^3

1 L = 1 dm^3

$1 m^3 = 100 (cm)^3 = 10^3 L$

Measurement of Mass

Mass is the quantity of matter contained in the sample and for the given sample it is constant and it does not depend upon the place.

weight is the force with with which the body is attracted towards the earth. It depends upon the acceleration due to gravity which varies from place to place. The mass of a substance can be determined very accurately in the laboratory by using analytical balance or electrical balance.

S.I. unit of mass is kilogram

1 Kg =1000 g

Units of Length

The S. I. unit of length is metre. It is also expressed in angstroms or nanometres or picometres.

Commonly used physical quantities and their derived units

Physical quantities and their derived units

Avogadro's hypothesis

Berzelius a Swedish chemist, gave a hypothesis called Berzelius hypothesis which states that :

Equal volume of all gases under similar conditions of temperature and pressure contain equal number of atoms.

For Example : Hydrogen + Chlorine —————-> 2 Hydrogen chloride gas

1 vol   1 vol   2 vol

n atoms   n atoms   2n molecule

On dividing throughout by 2n

½ a

This implies that one compound atom of hydrogen chloride is made up of ½ atom of hydrogen and ½ atom of chlorine. This is in direct conflict with Dalton's atomic theory which states that atoms are the ultimate particles of elements and are indivisible. This hypothesis was therefore rejected.

Avogadro's hypothesis

It states that equal volume of all gases under similar condition of temperature and pressure contain equal number of molecules .

Applications of Avogadro's law

1) In the calculation of atomicity of Elementary gases :

Atomicity of an elementary substance is defined as the number of atoms of the element present in 1 molecule of the substance.

For example : Atomicity of oxygen is 2 while that of ozone is 3.

Hydrogen   +   oxygen   —

2n molecules   n molecules   2n molecules

On dividing throughout by 2n

1 molecule   ½ molecule   1 molecule

Thus one molecule of water contains ½ molecule of oxygen. But 1 molecule of water contains 1 atom of oxygen.

Hence ½ molecule = 1 molecule of oxygen

1 molecule of oxygen= 2 atoms of oxygen = 1 atom of oxygen= 2

2) To find the relationship between molecular mass and vapour density of a gas

Molecular mass = 2 × vapour density

3) To find the relationship between mass and volume of a gas 22.4 litres of any gas at STP weigh equal to the molecular mass of the gas expressed in grams. This is called gram molecular volume law.

Atomic and Molecular mass

As an atom is so small a particle that it cannot be seen or isolated ,therefore it is impossible to determine the actual mass of a single atom by weighing it.

The problem was finally solved by Avogadro's hypothesis. If equal volumes of two different gases are taken under similar conditions of temperature and pressure and then weighted, the ratio of their masses will be equal to the ratio of their single molecules. Thus, though the actual masses of the atoms could not be determined but their relative masses could be determined. If the atomic mass of the hydrogen is taken is 1 ,the relative atomic mass of oxygen is 16 .

Initially, the atomic masses of all the elements were obtained by comparing with the mass of hydrogen taken as 1 but by doing so, the atomic masses of most of the elements came out to be fractional.

Therefore carbon is taken reference for the determination of atomic masses.

Atomic mass of an element is the number of times an atom of that element is heavier than an atom of carbon taken as 12.

One atomic mass unit is equal to one twelfth of the mass of an atom of carbon 12 isotope.

The atomic mass of an element is the average relative mass of its atoms as compared with an atom of carbon 12 taken as 12.

Fractional abundance of an isotope is the fraction of the total number of atoms that is comprised of that particular isotope.

Fractional abundance

Gram Atomic mass

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

For Ex: Atomic mass of oxygen is 16 amu.

Gram atomic mass of oxygen = 16 g.

Molecular mass

The molecular mass of a substance is the number of times the molecule of the substance is

heavier than one twelfth the mass of an atom of carbon -12.

or

The molecular mass is equal to sum of its atomic masses of all the atoms present in one molecule of substance.

For Example : Mass of H= 1 u

Atomic mass of O =16 u

Molecular mass of water =2 × atomic mass of H+1 × atomic mass of O

=2 × 1 + 16 × 1

= 18 u

Gram molecular mass

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

For Example : Molecular mass of oxygen = 32u

Gram molecular mass of oxygen = 32 g

Mole Concept

1 Avogadro's number or Avogadro's constant (NA)

2 Mole

3 Importance of Avogadro's number and Mole Concept

Avogadro's number or Avogadro's constant (NA)

One gram atom of any element contains the same number of atoms and one gram molecule of any substance contains the same number of molecules.

The value was found to be 6.022137×10^{23}

The value generally used is 6.022×10^{23} .

This is called Avogadro's number or Avogadro's constant (NA)

Avogadro's number may be defined as the number of atoms present in one gram atom of the element or the number of molecules present in one gram molecule of the substance.

A mole is a chemist unit of counting particles such as atom, molecules, ions, electrons, protons which represent a value of 6.022×10^{23}

A mole of hydrogen atom means 6.022×10^{23} atoms of hydrogen whereas a mole of hydrogen molecule means 6.022×10^{23} molecules of hydrogen or $2 \times 6.022 \times 10^{23}$ atoms of hydrogen .

A mole of oxygen molecule means 6.022×10^{23} molecules of oxygen or $2 \times 6.022 \times 10^{23}$ atoms of oxygen.

A mole is defined as that amount of substance which has mass equal to gram atomic mass if the substance is atomic or gram molecular mass if the substance is molecular.

1 mole of carbon atoms =12 grams

1 mole of sodium atoms = 23 grams

1 mole of Oxygen atom = 16 grams

1 mole of Oxygen molecule = 32 grams

1 mole of water molecule = 18 grams

1 mole of carbon dioxide molecule = 44 grams

Mole

A mole is defined as that amount of substance which contains Avogadro's number of atoms if the substance is atomic or Avogadro's number of molecules if the substance is molecular.

1 mole of carbon atoms = 6.022×10^{23} atoms of carbon.

1 mole of sodium atom = 6.022×10^{23} atoms of sodium

1 mole of Oxygen atom = 6.022×10^{23} atoms of oxygen

1 mole of Oxygen molecule = 6.022×10^{23} molecules of oxygen

1 mole of water = 6.022×10^{23} molecules of water

In case of gases, a mole is defined as that amount of the gas which has a volume of 22.4 litres at STP.

1 mole of Oxygen gas = 22.4 litres of oxygen at STP

one mole of carbon dioxide gas = 22.4 litres of carbon dioxide at STP

A mole of an ionic compound is defined as that amount of the substance which has mass equal

to gram formula mass or which contains Avogadro's number of formula unit.

1 mole of NaCl = 58.5 grams of NaCl

1 mole of NaCl = 6.022 ×10^{23} formula units of NaCl = 6.022 ×10^{23} Na+ ion and 6.022 ×10^{23}

Cl– ion.

Importance of Avogadro's number and Mole Concept

In the calculation of actual mass of a single atom of an element or a single molecule of a substance.

In the calculation of the number of atoms or molecules in a given mass of the element of the compound.

In the calculation of the number of molecules present in a given volume of the gas under given conditions.

In the calculation of the size of the individual atoms and molecules assuming them to the spherical.

In the calculation of actual masses of 1 amu or 1 u.

1 amu = 1.6606 × 10^{-24} g

1 amu = 1.6606 × 10^{-27} kg

Empirical and Molecular Formula

Calculation of percentage composition

The percentage 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.

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

Step 2 : Calculate the percentage of elements or the constituents by applying the relation:

Percentage composition

Question : Calculate the percentage of Carbon, Hydrogen and oxygen in Ethanol?

Answer : mass % C $= \frac{24.022\ g}{46.069\ g} \times 100\ \%\ = 52.144\ \%$

Mass % H $= \dfrac{6.048\ g}{46.069\ g} \times 100\ \%\ = 13.13\%$

Mass % O $= \dfrac{15.999\ g}{46.069\ g} \times 100\ \%\ = 34.72\ \%$

Empirical and Molecular Formula

The empirical formula of a compound is the chemical formula which expresses the simplest whole number ratio of the atoms of the various elements present in one molecule of the compound.

For Ex: The empirical formula of benzene is CH, hydrogen peroxide is HO, Glucose is CH2O.

The empirical formula represents only the atomic ratio of the various elements present in its molecule.

The molecular Formula of a compound is the chemical formula which represents the true formula of its molecules. It expresses the actual number of atoms of various elements present in one molecule of the compound.

The molecular formula of benzene is C6H6 , hydrogen peroxide is H2O2 , Glucose is C6H6.

Molecular Formula = n × Empirical formula

where n is any integer such as 1, 2, 3 ,4 etc.

The value of n can be obtained from the relation:

calculation of n

Molecular mass = 2 × Vapour Density

Calculation of Empirical Formula

Step 1 : Convert the mass percentage into grams.

Step 2 : Calculate the number of moles.

number of moles

Step 3 : Calculate the simplest molar ratio: Divide the moles obtained in step 1 by the smallest quotient or the least value from amongst the values obtained for each element.

 Step 4 : Calculate the simplest whole number ratio.

Step 5 : Write the empirical formula.

Empirical Formula

The empirical Formula is C_2H_4O.

Calculation of molecular Formula

The molecular formula of a compound can de deduced from:

1) Empirical Formula

2) Molecular mass

Step 1 Calculation of the empirical formula from the percentage composition.

Step 2 Calculation of empirical formula mass by adding the atomic mass of all the atoms present in the empirical formula.

Step 3 Determination of the molecular mass of the compound from the given data.

Step 4 Determination of the value of n.

Step 5 Determination of the molecular formula

Question A compound contains 4.07 % hydrogen, 24.27 % carbon and 71.65 % chlorine.Its molar mass is 98.96 g.What are its empirical and molecular formula.

Answer

Element	Symbol	Percentage Of element	Atomic mass	Moles of element	Simplest Molar ratio	Simplest Whole no. Molar ratio
Carbon	C	24.27	12	2.02	1	1
Hydrogen	H	4.07	1	4.07	2	2
Chlorine	Cl	71.65	35.5	2.02	1	1

The empirical formula of the compound is CH_2Cl

Empirical mass of CH_2Cl = 12 + 2 ×1 +35.5 = 49.5

n = 2

Molecular formula = n × E. F.

= 2 × CH_2Cl

= $C_2H_4Cl_2$

Balancing Of A Chemical Equation

Balancing of a chemical equation means making the number of atoms of each element equal on both sides of the equation.

The methods of balancing equation are:

1) Hit and Trial Method:

The simplest method to balance a chemical equation is by hit and trial method.

Step 1 : Write down the correct formula of the reactants and products with plus sign in between with an arrow pointing from reactants to Products. This is called skeletal equation.

Step 2: Select the biggest formula from the Skeleton equation and equalise the number of atoms of each of its constituent elements on both sides of the chemical equation by suitable multiplication .

 Step 3: When an elementary gas appear as a reactant or a product ,the equation is balanced more easily by keeping the elementary gas in the atomic state. The balanced atomic equation is then made molecular by multiplying the whole equation by 2.

Question : Magnetic oxide when heated with hydrogen is reduced to iron and water is also produced. Write balanced equation for the reaction?

Answer : Magnetic oxide + Hydrogen ———-> Iron + Water

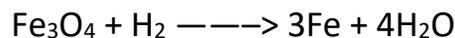$Fe_3O_4 + H_2$ ——–> $Fe + H_2O$

To equalise the number of atoms of Fe on both sides, multiply Fe by 3 , we have

$Fe_3O_4 + H_2$ ——–> $3Fe + H_2O$

The above equation has 4 atoms of O on L.H.S. and 1 atoms of O on R.H.S.

To equalise, multiply water by 4.

$Fe_3O_4 + H_2 ———> 3Fe + 4H_2O$

The above equation has 8 H atoms on R.H.S. and 2 H atoms on L.H.S.

To equalise ,multiply H_2 on L.H.S. by 4

$Fe_3O_4 + 4H_2 ———> 3Fe + 4H_2O$
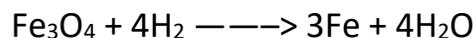
This a balanced chemical equation
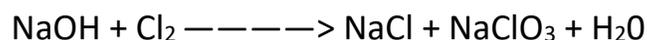
2) Partial Equation method

Step 1 The chemical reaction represented by the equation is supposed to proceed in two or more steps.

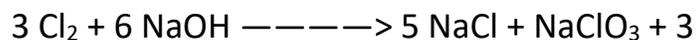Step 2 The Skeleton equation representing each step are written and then balanced by hit and trial method. These equations are known as partial equations.

Step 3 If necessary the partial equation are multiplied by suitable integers so as to cancel those intermediate products which do not occur in the final reaction .

Step 4 The partial equations are added up to get the final balanced equation.

Question : The skeleton equation to represent the action of chlorine on a hot solution of sodium hydroxide is

$NaOH + Cl_2 ————> NaCl + NaClO_3 + H_2O$

Balance this equation by the method of partial equations

Answer : $Cl_2 + H_2O ———-> H_2O +$ HC

$NaOH + HCl ————> NaCl + H_2O$

$NaOH + HClO ———-> NaClO + H_2O$

$3 NaClO ———-> 2 NaCl + NaClO_3$

$3 Cl_2 + 6 NaOH ————> 5 NaCl + NaClO_3 + 3$ H_2
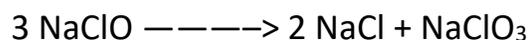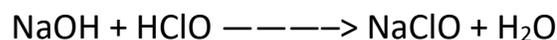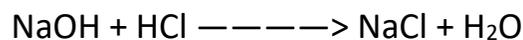
Summary-

1. **Matter**: Anything that occupies space and has mass.

2. **Element**: A pure substance which can neither be decomposed into nor built from simpler substances by any physical or chemical method. It contains only one kind of atoms.

3. **Compound**: A pure substance which can be decomposed into simpler substances by some suitable chemical method. It contains only one kind of molecules.
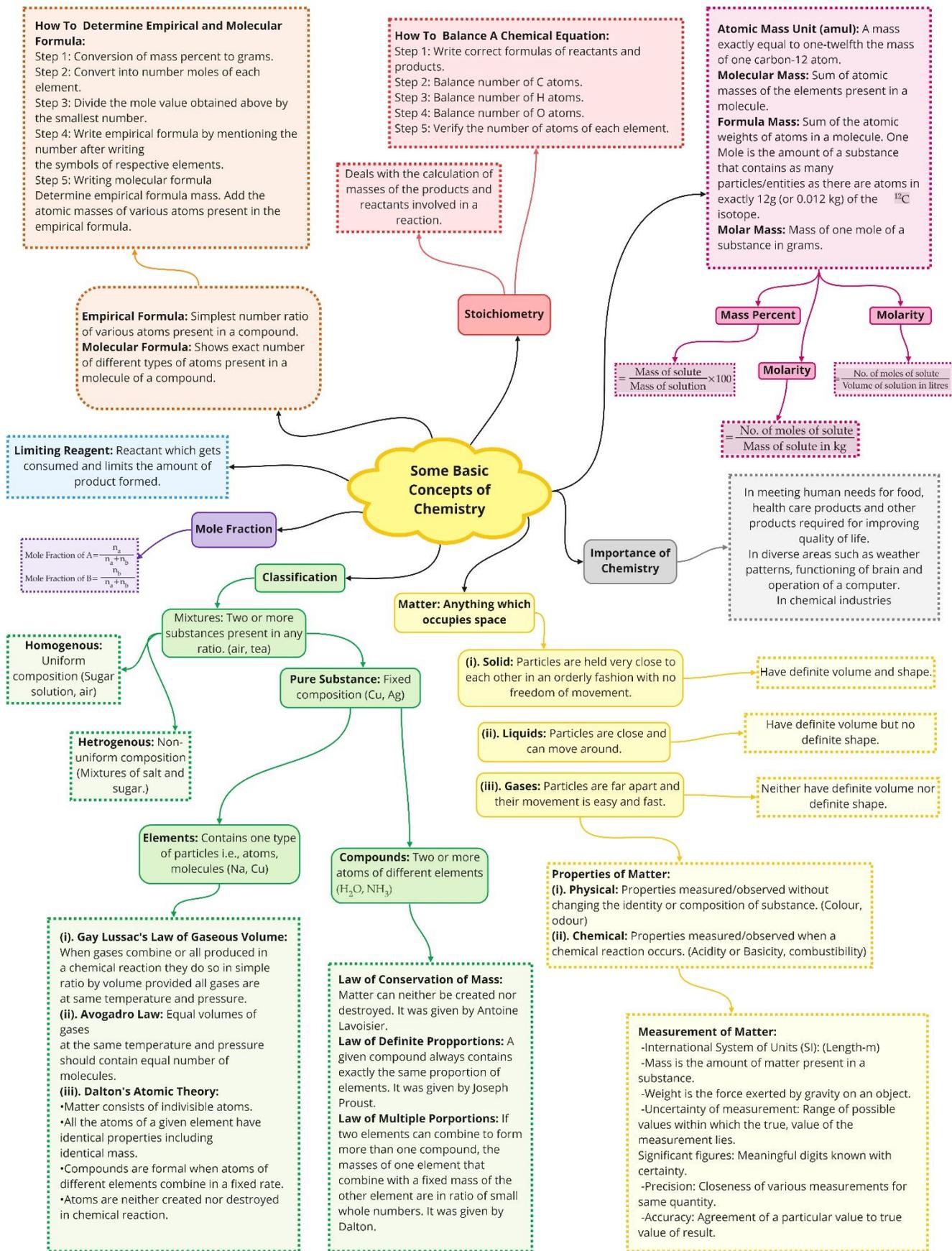
4. **Mixture**: A substance obtained by simple mixing of two or more pure substances.

5. **Law of Conservation of Mass**: During any physical or chemical change total mass of the products formed is equal to the total mass of the reactants consumed.

6. **Law of Constant Composition**: A chemical compound always contains same elements combined together in same proportion of mass.

7. **Law of Multiple Proportions**: When two elements combine with each other to form two or more than two compounds then the masses of one of the elements that combine with the fixed mass of the other, bear a simple whole number ratio to one another.

8. **Gay Lussac's law**: When gases react with each other they do so in volumes which bear a simple whole number ratio to one another and to the volumes of products, if there are also gases, provided all volumes are measured under similar conditions of temperature and pressure.

9. **Avogadro's Law**: Equal volume of all gases under similar conditions contain equal number of molecules.

10. **Atom**: The smallest particle of an element that takes part in chemical reactions.

11. **Molecule**: The smallest particle of a substance that has independent existence.

12. **Atomicity**: The number of atoms in a molecule of the elementary substance.

13. **Unified Mass (u)**: One-twelfth of the actual mass of an atom of carbon (C-12). It is equal to 1.66×10^{-27} kg.

14. **Atomic Mass**: The average relative mass of an atom of the element as compared with mass of a carbon atom (C-12) taken as 12 u.

15. **Molecular Mass**: The average relative mass of a molecule of the substance as compared with mass of an atom of carbon (C-12) taken as 12 u.

16. **Gram Atomic Mass**: The mass of 1 mole of atoms (6×10^{23}) in g is called gram atomic mass.

17. **Gram Molecular Mass**: The mass of 1 gram molecule of compound expressed in grams.

18. **Avogadro's Number (NA)**: 6.022×10^{23}.

19. **Mole**: 6.022×10^{23} specified particles.

20. **Molar Mass**: Mass of one mole particles of the substance.

21. **Gram Molecular Volume (G.M.V.)**: Volume occupied by one mole molecules of the gaseous substance. Its value is equal to 22.4 L and S.T.P.

22. **Empirical Formula**: The formula which gives the simplest whole number ratio of atoms of different elements present in the molecule of the compound. Molecular formula is whole number multiple of empirical formula.

23. **Molarity (M)**: Number of moles of solute per litre of solution. Expressed as moles per litre or moles per dm3 or Molar (M).

24. Molarity changes with change in temperature because volume of the solution changes with change in temperature.

25. K = °C + 273.15

26. $°F = \dfrac{9}{5} (°C) + 32$

Class : 11th Chemistry
Chapter- 1: Some Basic Concepts of Chemistry

How To Determine Empirical and Molecular Formula:
Step 1: Conversion of mass percent to grams.
Step 2: Convert into number moles of each element.
Step 3: Divide the mole value obtained above by the smallest number.
Step 4: Write empirical formula by mentioning the number after writing
the symbols of respective elements.
Step 5: Writing molecular formula
Determine empirical formula mass. Add the atomic masses of various atoms present in the empirical formula.

How To Balance A Chemical Equation:
Step 1: Write correct formulas of reactants and products.
Step 2: Balance number of C atoms.
Step 3: Balance number of H atoms.
Step 4: Balance number of O atoms.
Step 5: Verify the number of atoms of each element.

Atomic Mass Unit (amul): A mass exactly equal to one-twelfth the mass of one carbon-12 atom.
Molecular Mass: Sum of atomic masses of the elements present in a molecule.
Formula Mass: Sum of the atomic weights of atoms in a molecule. One Mole is the amount of a substance that contains as many particles/entities as there are atoms in exactly 12g (or 0.012 kg) of the ^{12}C isotope.
Molar Mass: Mass of one mole of a substance in grams.

Deals with the calculation of masses of the products and reactants involved in a reaction.

Empirical Formula: Simplest number ratio of various atoms present in a compound.
Molecular Formula: Shows exact number of different types of atoms present in a molecule of a compound.

Stoichiometry

Mass Percent

Molarity

$$= \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100$$

Molarity

$$= \frac{\text{No. of moles of solute}}{\text{Volume of solution in litres}}$$

$$= \frac{\text{No. of moles of solute}}{\text{Mass of solute in kg}}$$

Limiting Reagent: Reactant which gets consumed and limits the amount of product formed.

Mole Fraction

$$\text{Mole Fraction of A} = \frac{n_a}{n_a + n_b}$$
$$\text{Mole Fraction of B} = \frac{n_b}{n_a + n_b}$$

Some Basic Concepts of Chemistry

Importance of Chemistry

In meeting human needs for food, health care products and other products required for improving quality of life.
In diverse areas such as weather patterns, functioning of brain and operation of a computer.
In chemical industries

Classification

Matter: Anything which occupies space

Mixtures: Two or more substances present in any ratio. (air, tea)

Pure Substance: Fixed composition (Cu, Ag)

Homogenous: Uniform composition (Sugar solution, air)

Hetrogenous: Non-uniform composition (Mixtures of salt and sugar.)

(i). Solid: Particles are held very close to each other in an orderly fashion with no freedom of movement.

Have definite volume and shape.

(ii). Liquids: Particles are close and can move around.

Have definite volume but no definite shape.

(iii). Gases: Particles are far apart and their movement is easy and fast.

Neither have definite volume nor definite shape.

Elements: Contains one type of particles i.e., atoms, molecules (Na, Cu)

Compounds: Two or more atoms of different elements (H_2O, NH_3)

Properties of Matter:
(i). Physical: Properties measured/observed without changing the identity or composition of substance. (Colour, odour)
(ii). Chemical: Properties measured/observed when a chemical reaction occurs. (Acidity or Basicity, combustibility)

(i). Gay Lussac's Law of Gaseous Volume:
When gases combine or all produced in a chemical reaction they do so in simple ratio by volume provided all gases are at same temperature and pressure.
(ii). Avogadro Law: Equal volumes of gases
at the same temperature and pressure should contain equal number of molecules.
(iii). Dalton's Atomic Theory:
•Matter consists of indivisible atoms.
•All the atoms of a given element have identical properties including identical mass.
•Compounds are formal when atoms of different elements combine in a fixed rate.
•Atoms are neither created nor destroyed in chemical reaction.

Law of Conservation of Mass:
Matter can neither be created nor destroyed. It was given by Antoine Lavoisier.
Law of Definite Propportions: A given compound always contains exactly the same proportion of elements. It was given by Joseph Proust.
Law of Multiple Porportions: 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 numbers. It was given by Dalton.

Measurement of Matter:
-International System of Units (SI): (Length-m)
-Mass is the amount of matter present in a substance.
-Weight is the force exerted by gravity on an object.
-Uncertainty of measurement: Range of possible values within which the true, value of the measurement lies.
Significant figures: Meaningful digits known with certainty.
-Precision: Closeness of various measurements for same quantity.
-Accuracy: Agreement of a particular value to true value of result.

Important Questions

Multiple Choice questions-

Question 1. Formula of Ferric Sulphate is:

(a) $FeSo_4$

(b) $Fe(So_4)_3$

(c) $Fe_2(So_4)_3$

(d) Fe_2So_4

Question 2. Approximate atomic weight of an element is 26.89. If its equivalent weight is 8.9, the exact atomic weight of element would be

(a) 26.89

(b) 8.9

(c) 17.8

(d) 26.7

Question 3. The total number of atoms represented by the compound $CuSO_4 . 5H_2O$ is

(a) 27

(b) 21

(c) 5

(d) 8

Question 4. An atom is 10 times heavier than 1/12th of mass of a carbon atom (C – 12). The mass of the atom in a.m.u. is

(a) 10

(b) 120

(c) 1.2

(d) 12

Question 5. 81.4 g sample of ethyl alcohol contains 0.002 g of water. The amount of pure ethyl alcohol to the proper number of significant figures is

(a) 81.398 g

(b) 71.40 g

(c) 91.4 g

(d) 81 g

Question 6. Which of the following halogen can be purified by sublimation?

(a) F_2

(b) Cl_2

(c) Br_2

(d) I_2

Question 7. 1 mol of CH_4 contains

(a) 6.02×10^{23} atoms of H

(b) 4 g atom of Hydrogen

(c) 1.81×10^{23} molecules of CH_4

(d) 3.0 g of carbon

Question 8. The prefix zepto stands for

(a) 10_9

(b) 10^{-12}

(c) 10^{-15}

(d) 10^{-21}

Question 9. Which has maximum number of atoms?

(a) 24 g of C (12)

(b) 56 g of Fe (56)

(c) 27 gof Al (27)

(d) 108 g of Ag (108)

Question 10: Irrespective of the source, pure sample, of water always yields 88.89% mass of oxygen and 11.11% mass of hydrogen. This is explained by the law of

(a) Conservation of Mass

(b) Multiple Proportions

(c) Constant Composition

(d) Constant Volume

Question 11. Hemoglobin contains 0.33% of iron by weight. The molecular weight of hemoglobin is approximately 67200. The number of iron atoms (At. wt. of Fe = 56) present in one molecule of hemoglobin is

(a) 6

(b) 1

(c) 4

(d) 2

Question 12. The -ve charged particles is called:

(a) Anion

(b) Cation

(c) Radical

(d) Atom

Question 13. Which of the following contains same number of carbon atoms as are in 6.0 g of carbon (C – 12)?

(a) 6.0 g Ethane

(b) 8.0 g Methane

(c) 21.0 g Propane

(d) 28.0 g CO

Question 14. The density of a gas is 1.78 gL^{-1} at STP. The weight of one mole of gas is

(a) 39.9 g

(b) 22.4 g

(c) 3.56 g

(d) 29 g

Question 15. Molarity of 0.2 N H_2SO_4 is

(a) 0.2

(b) 0.4

(c) 0.6

(d) 0.1

Very Short:

1. What is chemistry?

2. How has chemistry contributed towards nation's development?

3. Differentiate solids, liquids & gases in terms of volume & shapes.

4. Name the different methods that can be used for separation of components of a mixture.

5. Classify following as pure substances and mixtures – Air, glucose, gold, odium and milk

6. What is the difference between molecules and compounds? Give examples of each.?

7. How can we separate the components of a compound?

8. How are physical properties different from chemical properties?

9. What are the two different system of measurement?

10. What is the SI unit of density?

Short Questions:

1. Define Mole. What is its numerical value?

2. Define molarity. Is it affected by a change in temperature?

3. What do you mean by Precision and accuracy?

4. Distinguish between fundamental and the derived units.

5. Define molality and write its temperature dependence.

6. Distinguish between an atom and a molecule.

7. Derive the SI unit of Joule (J) in terms of fundamental units.

Long Questions:

1. State the law of Multiple Proportions. Explain with two examples.

2. State the law of Constant Composition. Illustrate with two examples.

3. Define empirical formula and molecular formula. How will you establish a relationship between the two? Give examples.

4. In the commercial manufacture of nitric acid, how many moles of NO2 produce 7.33 mol HN03 in the reaction

   $3 NO_2(g) + H_2O(1) \rightarrow 2HNO_3(aq) + NO(g)$

5. A sample of $NaNO_3$ weighing 0.83 g is placed in a 50,0 mL volumetric flask. The flask is then filled with water upon the etched mark. What is the molarity of the solution?

Assertion Reason Questions:

1. In the following questions, a statement of Assertion (A) followed by a statement of Reason (R) is given. Choose the correct option out of the choices given below each question.

   Assertion (A) : The empirical mass of ethene is half of its molecular mass.

   Reason (R) : The empirical formula represents the simplest whole-number

   ratio of various atoms present in a compound.

   (i) Both A and R are true and R is the correct explanation of A.

   (ii) A is true but R is false.

(iii) A is false but R is true.

(iv) Both A and R are false.

2. In the following questions, a statement of Assertion (A) followed by a statement of Reason (R) is given. Choose the correct option out of the choices given below each question.

Assertion (A) : One atomic mass unit is defined as one twelfth of the mass of

one carbon-12 atom.

Reason (R) : Carbon-12 isotope is the most abundunt isotope of carbon

and has been chosen as standard.

(i) Both A and R are true and R is the correct explanation of A.

(ii) Both A and R are true but R is not the correct explanation of A.

(iii) A is true but R is false.

(iv) Both A and R are false.

Case Study Based Question:

1. Chemistry is the science of molecules and their transformations. It is the science not so much of the one hundred elements but of the infinite variety of molecules that may be built from them. Chemistry plays a central role in science and is often intertwined with other branches ofscience.to understand the basic concepts of chemistry, which begin with the concept of matter. Let us start with the nature of matter. matter can exist in three physical states viz. solid, liquid and gas. Particles are held very close to each other in solids in an orderly fashion and there is not much freedom of movement. In liquids, the particles are close to each other but they can move around. However, in gases, the particles are far apart as compared to those present in solid or liquid states and their movement is easy and fast. different states of matter exhibit the following characteristics:

I. Solids have definite volume and definite shape.

II. Liquids have definite volume but do not have definite shape. They take the shape of the container in which they are placed.

II. Gases have neither definite volume nor definite shape. They completely occupy the space in the container in which they are placed.

Matter can be classified as mixture or pure substance. A mixture may be homogeneous or heterogeneous. Pure substances can further be classified into elements and compounds. Particles of an element consist of only one type of atoms. These particles may exist as atoms or molecules. When two or more atoms of different elements combine together in a definite ratio, the molecule of a compound is obtained.

Every substance has unique or characteristic properties. These properties can be classified into two categories — physical properties, such as colour, odour, melting point, boiling point, density, etc., and chemical properties, like composition, combustibility, reactivity with acids and bases, etc. Physical properties can be measured or observed without changing the identity or the composition of the substance. The measurement or observation of chemical properties requires a chemical change to occur. Measurement of physical properties does not require occurrence of a chemical change.

(1) Which of the following state of matter have definite volume but do not have definite shape ?

(a) Solid

(b) Liquid

(c) Gas

(d) Plasma

(2) Particles are held very close to each other in … in an orderly fashion and there is not much freedom of movement.

(a) Liquid

(b) Gas

(c) Solid

(d) Plasma

(3) Particles of …. consist of only one type of atom.

(a) Compound

(b) Mixture

(c) Element

(d) All the above

(4) Water molecule comprises …hydrogen atoms and … oxygen atom.

(a) One , two

(b) Three , one

(c) One , three

(d) Two , one

(5) Which of the following is not an example of Physical Properties of substance.?

Hydrogen + Oxygen → Hydrogen Peroxide

2g   32g   34g

Here, the masses of oxygen (i.e., 16 g and 32 g), which combine with a fixed mass of hydrogen (2g) bear a simple ratio, i.e., 16:32 or 1:2.

4) **Gay Loussac's Law of Gaseous Volumes-** This law was given by Gay Loussac in 1808. He observed that when gases combine or are produced in a chemical reaction they do so in a simple ratio by volume, provided all gases are at the same temperature and pressure.

5) **Avogadro's Law-** In 1811, Avogadro proposed that equal volumes of all gases at the same temperature and pressure should contain equal number of molecules.

In 1808, Dalton published 'A New System of Chemical Philosophy', in which he proposed the following :

(1) Matter consists of indivisible atoms.

(2) All atoms of a given element have identical properties, including identical mass. Atoms of different elements differ in mass.

(3) Compounds are formed when atoms of different elements combine in a fixed ratio.

(4) Chemical reactions involve reorganisation of atoms. These are neither created nor destroyed in a chemical reaction.

1) ………… refers to the closeness of various measurements for the same quantity.

  a) Accuracy

  b) Reliability

  c) Precision

  d) Uncertainty

2) Law of Conservation of mass was put forth by ………. in 1789.

  a) Joseph Proust

  b) Antoine Lavoisier

  c) Joseph Louis

  d) Gay Loussac

3) Which of the following number has two significant figures.

  a) 00052

b) 052

c) 52

d) 0.0052

4) ……….. is the agreement of a particular value to the true value of the result.

   a) Accuracy

   b) Reliability

   c) Precision

   d) Uncertainty

5) Law of Multiple Proportions proposed by ..

   a) Joseph Proust

   b) Antoine Lavoisier

   c) Joseph Louis

   d) John Dalton

Answer Key:

MCQ

1. (c) $Fe_2 (SO_4)_3$

2. (a) 26.89

3. (b) 21

4. (a) 10

5. (a) 81.398 g

6. (d) I_2

7. (c) 4 g atom of Hydrogen

8. (d) 10^{-21}

9. (a) 24 g of C (12)

10. (c) Constant Composition

11. (c) 4

12. (a) Anion

13. (b) 8.0 g Methane

14. (a) 39.9 g

15. (d) 0.1

Very Short Answer:

1. Chemistry is the branch of science that studies the composition, properties and interaction of matter.

2. chemical principles are important in diverse areas such as weather patterns, functioning of brain, operation of a computer, chemical industries, manufacturing , fertilizers, alkalis, acids, salts, dyes, polymers, drugs, soaps, detergents, metals, alloys, contribute in a big way to national economy.

3.

roperty	Solids	Liquids	Gases
1. Volume	Definite	Definite	**Not definite**
2. Shape	**Fixed**	**Not fixed, take the shape of container,**	**Not fixed, takes the shape of the container**

4. The components of a mixture can be separated by physical methods like handpicking, filtrations, crystallization, distillation etc.

5.

Pure Substances	Mixtures
Glucose	Air
Gold	Milk
Sodium	

6. Molecules consist of different atoms or same atoms. e.g. molecule of hydrogen contains two atoms of hydrogen whereas molecule of water contain two atoms of hydrogen and one of oxygen.

Compound is formed when two or more than two different atoms combine in fire propo e.g. water –ration carbon dioxide, sugar etc.

7. The constituents of a compound cannot be separated by physical methods. They can only be separate by chemical methods.

8. Physical properties are those properties which can be measured or observed without changing the identity or the composition of the substance whereas the measurement of chemical properties require a chemical change to occur e.g. color, odour etc. are physical properties and combustion, basicity etc. are chemical properties.

9. The different system of measurement are English system and the metric system.

10. The SI Unit of density is Kg m^{-3} or kg/m^3

Short Answer:

Ans: 1. A mole is the amount of a substance that contains as many entities (atoms, molecules, or other particles) as there are atoms in exactly 0.012 kg or 12 g of the carbon-12 isotope.

Ans: 2. The molarity of a solution is defined as the number of moles of the solute present per liter of the solution. It is represented by the symbol M. Its value changes with the change in temperature.

Ans: 3. Precision and accuracy: The term precision refers to the closeness of the set of values obtained from identical measurements of a quantity.

Accuracy refers to the closeness of a single measurement to its true value.

Ans: 4. Fundamental units: Fundamental units are those units by which other physical units can be derived. These are mass (M), Length (L), time (T), temperature (°).

Derived units: The units which are obtained by the combination of the fundamental units are called derived units.

Ans: 5. Molality (m) = $\dfrac{\text{Mole of solute}}{\text{Mass of the solvent in kg}}$

The molality of the solution does not depend upon the temperature.

Ans: 6. Atom: An atom is the smallest particle of an element that takes part in a chemical reaction. It may or may not be capable of independent existence.

Molecule: It is the smallest particle of a substance (element or compound) that is capable of independent existence

Ans: 7. Joule is the SI unit of work or energy

As work $=$ force $\times$ distance

$\qquad = $ (mass $\times$ acceleration) $\times$ distance

$\qquad = $ Mass $\times \dfrac{\text{distance}}{\text{time}^2} \times$ distance

$\qquad = \dfrac{\text{mass} \times (\text{distance})^2}{\text{time}^2}$

Hence $\qquad J = \dfrac{\text{kg} \times \text{m}^2}{\text{s}^2} = \text{kg m}^2\text{s}^{-2}.$

Long Answer:

Ans: 1. The Law of Multiple Proportions states:

"When two elements combine to form two or more than two chemical compounds than the weights of one of elements which combine with a fixed weight of the other, bear a simple ratio to one another.

Examples:

1. Compound of Carbon and Oxygen: C and O combine to form two compounds CO and CO_2.

In CO_2 12 parts of wt. of C combined with 16 parts by wt. O.

In CO_2 12 parts of wt. of C combined with 32 parts by wt. of O.

If the weight of C is fixed at 12 parts by wt., then the ratio in the weights of oxygen which combine with the fixed wt. of C (= 12) is 16: 32 or 1: 2.

Thus, the weight of oxygen bears a simple ratio of 1: 2 to each other.

2. Compounds of Sulphate (S) and Oxygen (O):

S forms two oxides with O, viz., SO_2 and SO_3

In SO_2, 32 parts of wt. of S combine with 32 parts by wt. of O.
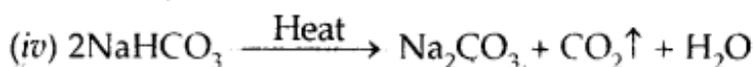
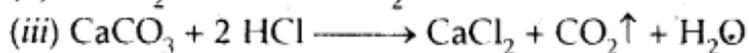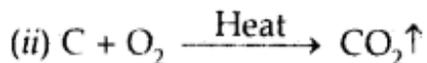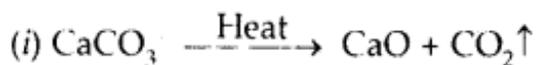In SO_3, 32 parts of wt. of S combine with 48 parts by wt. of O.

If the wt. of S is fixed at 32 parts, then' the ratio in the weights of oxygen which combine with the fixed wt. of S is 32: 48 or 2: 3.

Thus, the weights of oxygen bear a simple ratio of 2: 3 to each other.

Ans: 2. Law of Constant Composition of Definite Proportions states: "A chemical compound is always found, to be made up of the same elements combined together in the same fixed proportion by weight".

Examples:

1. CO_2 may be prepared in the laboratory as follows:

(i) $CaCO_3 \xrightarrow{\text{Heat}} CaO + CO_2\uparrow$

(ii) $C + O_2 \xrightarrow{\text{Heat}} CO_2\uparrow$

(iii) $CaCO_3 + 2\,HCl \longrightarrow CaCl_2 + CO_2\uparrow + H_2O$

(iv) $2NaHCO_3 \xrightarrow{\text{Heat}} Na_2CO_3 + CO_2\uparrow + H_2O$

In all the above examples, CO2 is made up of the same elements i. e., Carbon (C) and Oxygen (O) combined together in the same fixed proportion by weight of 12: 32 or 3: 8 by weight.

Ans: 3. The empirical formula of a compound expresses the simplest whole-number ratio of the atoms of the various elements present in one molecule of the compound.

For example, the empirical formula of benzene is CH and that of glucose is CH2O. This suggests that in the molecule of benzene one atom of Carbon (C) is present for every atom of Hydrogen (H). Similarly in the molecule of glucose (CH2O), for every one atom of C, there are two atoms of H and one atom of O present in its molecule. Thus, the empirical formula of a compound represents only the atomic ratio of various elements present in its molecule.

The molecular formula of a compound represents the true formula of its molecule. It expresses the actual number of atoms of various elements present in one molecule of a compound. For example, the molecular formula of benzene is C_6H_6 and that of glucose is $C_6H_{12}O_6$. This suggests that in one molecule of benzene, six atoms of C and 6 atoms of H are present. Similarly, one molecule of glucose ($C_6H_{12}O_6$) actually contains 6 atoms of C, 12 atoms of H, and 6 atoms of O.

Relation between the empirical and molecular formula

Molecular formula = n × Empirical formula where n is an integer such as 1, 2, 3…

When n = 1; Molecular formula = Empirical formula

When n = 2; Molecular formula = 2 × Empirical formula.

The value of n can be obtained from the relation.

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

The molecular mass of a volatile substance can be determined by Victor Meyer's method or by employing the relation.

Molecular mass = 2 × vapour density.

Empirical formula mass can however be obtained from its empirical formula simply by adding the atomic masses of the various atoms present in it.

Thus, the empirical formula mass of glucose CH_2O

= 1 × 12 + 2 × 1 + 1 × 16 = 30.0 u.

Ans: 4. 2 mols of HNO_3 are produced by 3 mols of NO_2

7.33 mol HNO_3 are produced by $\frac{3 \times 7.33}{2}$ mol of NO_2

= 10.995 mols.

Ans: 5. Molar mass of $NaNO_3$ = 23 + 14 + 3 × 16 = 85 g mol^{-1}

Molarity = Number of moles of solute Volume of solution in L

$$= \frac{0.83 \times 1000}{85 \times 50}$$

= 0.196 M.

Assertion Reason Answer:

1. (i) Both A and R are true and R is the correct explanation of A.

2. (ii) Both A and R are true but R is not the correct explanation of A.

Case Study Answer:

1. **Answer:**

   (1) (b) Liquid

   (2) (c) Solid

   (3) (c) Element

(4) (d) Two , one.

(5) (d) Composition

2. Answer:

(1) (c) Precision

(2) (b) Antoine Lavoisier

(3) (d) 0.0052

(4) (a) Accuracy

(5) (d) John Dalton