1. SOME BASIC CONCEPTS OF CHEMISTRY

Chemistry is the branch of Science that deals with the preparation, properties, structure and reactions of material substances.

Some important branches of Chemistry are:

1. Inorganic Chemistry
2. Organic Chemistry
3. Physical Chemistry
4. Analytical Chemistry
5. Polymer Chemistry
6. Biochemistry
7. Medicinal Chemistry
8. Industrial Chemistry
9. Hydrochemistry
10. Electrochemistry
11. Green Chemistry etc.

Matter: Matter is anything that occupies space and has a definite mass. Based on the physical state we can divide matter into different categories.

1. Solid state
2. Liquid state
3. Gaseous state
4. Plasma state
5. Bose-Einstein condensate
6. Fermionic condensate
7. Quark-Gluon Plasma

In earth crust, matter mainly exists in three physical states – solid state, liquid state and gaseous state.

In solids, the particles are orderly arranged and they are very close to each other. The particles cannot move freely. So solids have definite shape and definite volume.

In liquids, the particles are close to each other but they can move around. So, liquids have definite volume but do not have definite shape.

In gases, the particles are far apart as compared to those present in solid or liquid state and their movement is easy and fast. So they do not have definite shape and volume. They take the shape of the container in which they are placed. Also they occupy the complete space of the container in which they are placed.

The three states of matter are interconvertible by changing the conditions of temperature and pressure.

$$\text{Solid} \underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}} \text{liquid} \underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}} \text{Gas}$$

Classification of matter

Based on the chemical composition, matter can be divided into two categories – pure substances and mixtures.

Pure substances contain only one type of particles. E.g. Sodium (Na), Potassium (K), Hydrogen (H), Oxygen (O), Helium (He), carbon dioxide (CO_2), water (H_2O), ammonia (NH_3), cane sugar ($C_{12}H_{22}O_{11}$) etc.

Pure substances are classified into two – ***elements and compounds***.

Elements are pure substances which contain only one type of atom. The term element was first introduced by Robert Boyle, the father of ancient Chemistry. Now there are 118 elements starting from hydrogen ($_1H$) and ending in Oganesson ($_{118}Og$). Some elements exist as monoatomic, some are diatomic and some others are polyatomic. E.g. all metals (Sodium, Potassium, Calcium etc) and noble gases (Helium, Neon etc.) are monoatomic. Hydrogen, Nitrogen, Oxygen etc are diatomic. Phosphorus (P_4) and Sulphur (S_8) are polyatomic.

Compounds are pure substances which contain more than one type of atoms. They are formed by the combination of two or more atoms of different elements in a definite ratio. Their constituents cannot be separated by physical methods, but they can be separated by chemical methods.

E.g. CO_2, H_2O, NH_3, H_2SO_4 etc.

Mixtures contain more than one type of particles. The components of a mixture can be separated by using physical methods like filtration, crystallisation, distillation etc.

E.g. all types of solutions, gold ornaments, sea water, muddy water, air etc.

There are two types of mixtures – homogeneous and heterogeneous mixtures.

Mixtures having uniform composition throughout are called ***homogeneous mixtures***. Here the components are completely mixed with each other. E.g. all type of solutions, air etc.

Mixtures having different compositions at different parts are called ***heterogeneous mixtures***.

E.g. sea water, soil. Muddy water etc.

Physical and chemical properties

The properties or characteristics of matter can be classified into two types — ***physical properties and chemical properties.***

Properties which can be measured or observed without changing the composition or identity of the substance are called physical properties. Measurement of physical properties does not require the occurrence of a chemical change.

E.g. colour, odour, melting point, boiling point, density, mass etc.

Properties which can be measured only with the occurrence of a chemical change are called chemical properties. E.g. composition, combustibility, reactivity with acids and bases, etc.

Measurement of physical properties

Any quantitative observation or measurement is represented by a number followed by units in which it is measured. Earlier, two different systems of measurement were used: the English System and the Metric System.

Now a days, a common standard system known as International System of Units (SI) is used. This system has **seven base units** and they are length, mass, time, electric current, thermodynamic temperature, amount of substance and luminous intensity. Their SI units are as follows:

Physical Quantity	SI unit & its symbol
Length	metre (m)
Mass	kilogram (kg)
Time	second (s)
Electric current	ampere (A)
Thermodynamic temperature	kelvin (K)
Amount of substance	mole (mol)
Luminous intensity	candela (cd)

Mass and Weight

Mass is the amount of matter present in a body. It is a constant quantity. Its SI unit is kilogram (kg). Weight is the gravitational force acting on a body. It is a variable quantity. i.e. it changes with place. Its SI unit is newton (N).

Volume (V)

It is the amount of space occupied by a body. Its SI unit is m^3. In Chemistry, smaller volumes are used. Hence, volume is often denoted in cm^3, dm^3, mL, L etc.

$1\ m^3 = 10^6\ cm^3$ $\quad$ $1\ L = 10^3\ cm^3\ (mL)$ $\quad$ $1 cm^3 = 1\ mL$

$1\ dm^3 = 10^3\ cm^3$ $\quad$ $1\ dm^3 = 1\ L$

Density (d)

It is the amount of mass per unit volume.

i.e. density = mass/volume. Its SI unit is kg/m^3. But it is commonly expressed in g/cm^3.

Temperature (T)

It is the degree of hotness or coldness of a body. It is commonly expressed in degree celsius (^0C). Other units are degree fahrenheit (^0F), kelvin (K) etc. Its SI unit is kelvin (K).

Degree celsius and degree fahrenheit are related as:

$$^0F = \frac{9}{5}(^0C) + 32$$

Degree celsius and kelvin are related as:

$$K = {}^0C + 273.15 \qquad \text{OR,} \qquad K = {}^0C + 273$$

Precision and Accuracy

Precision refers to the closeness of various measurements for the same quantity. But, accuracy is the agreement of a particular value to the true value of the result.

Scientific Notation

It is an exponential notation in which a number is represented in the form $N \times 10^n$, where n is an exponent having positive or negative values and N is a number (called digit term) which varies between 1.000... and 9.999.....

While writing scientific notation, the value of the exponent 'n' becomes positive, when the decimal is shifted to left and it becomes negative, when the decimal is shifted to right.

E.g. the scientific notation of 368.9 is 3.689×10^2 and that of 0.000563 is 5.63×10^{-4}.

Significant Figures

Every experimental measurement has some amount of uncertainty associated with it. The uncertainty in the experimental or the calculated values is indicated by mentioning the number of significant figures. ***Significant figures are meaningful digits which are known with certainty***. The uncertainty is indicated by writing the certain digits and the last uncertain digit.

There are certain rules for determining the number of significant figures. These are:

1. All non-zero digits are significant. For example in 285 cm, there are three significant figures and in 0.25 mL, there are two significant figures.
2. Zeros preceding to first non-zero digit are not significant. Such zero indicates the position of the decimal point. Thus, 0.03 has one significant figure and 0.0052 has two significant figures.
3. Zeros between two non-zero digits are significant. Thus, 2.005 has four significant figures.

4. Zeros at the end or right of a number are significant if they are on the right side of the decimal point; otherwise, they are not significant. For example, 0.200 g has three significant figures.
5. Exact numbers have an infinite number of significant figures. For example, in 2 balls or 20 eggs, there are infinite significant figures since these are exact numbers and can be represented by writing infinite number of zeros after placing a decimal i.e., 2 = 2.000000 or 20 = 20.000000
6. In numbers written in scientific notation, all digits are significant. E.g. 4.01×10^2 has three significant figures, and 8.256×10^{-3} has four significant figures.

Rounding off a number to the required number of significant figures

The rules related to rounding off a number are:

1) If the rightmost digit to be removed is more than 5, the preceding number is increased by one. E.g. 1.386 can be round off to three significant figures, by removing 6. So it becomes 1.39.
2) If the rightmost digit to be removed is less than 5, the preceding number is not changed. E.g. In 4.334, if 4 is to be removed, then the result is rounded upto 4.33.
3) If the rightmost digit to be removed is 5, then the preceding number is not changed if it is an even number but it is increased by one if it is an odd number. For example, if 6.35 is to be rounded by removing 5, we have to increase 3 to 4 giving 6.4 as the result. But if 6.25 is to be rounded off, it becomes 6.2.

LAWS OF CHEMICAL COMBINATIONS

The combination of elements to form compounds is governed by the following five basic laws:

1. ***Law of Conservation of Mass (Law of indestructibility of matter):*** This law was proposed by **Antoine Lavoisier**. It states that matter can neither be created nor destroyed. Or, in a chemical reaction, the total mass of reactants is equal to the total mass of products. Chemical equations are balanced according to this law.

   Illustration

   Consider the reaction $2H_2 + O_2 \longrightarrow 2H_2O$

   Here 4 g of H_2 combines with 32 g of O_2 to form 36 g of water.

   Total mass of reactants = 4 + 32 = 36g

   Total mass of products = 36 g

2. **Law of Definite Proportions (Law of definite composition):** This law was proposed by **Joseph Proust**. It states that a given compound always contains exactly the same proportion of elements by weight. Or, the same compound always contains the same elements combined in a fixed ratio by mass.

Illustration: Carbon dioxide can be formed in the atmosphere by various methods like respiration, burning of fuels, reaction of metal carbonates and bicarbonates with acid etc. All these samples of CO_2 contain only two elements Carbon and Oxygen combined in a mass ratio 3:8.

3. **Law of Multiple Proportions**: This law was proposed by **John Dalton**. It states that if two elements can combine to form more than one compound, the different masses of one of the elements that combine with a fixed mass of the other element, are in small whole number ratio.

Illustration: Hydrogen combines with oxygen to form two compounds – water and hydrogen peroxide.

$$\underset{2g}{H_2} + \underset{16g}{\tfrac{1}{2} O_2} \longrightarrow \underset{18g}{H_2O}$$

$$\underset{2g}{H_2} + \underset{32g}{O_2} \longrightarrow \underset{34g}{H_2O_2}$$

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 Lussac's Law of Gaseous Volumes**: This law was proposed by **Gay Lussac**. It states that when gases combine to form gaseous products, their volumes are in simple whole number ratio at constant temperature and pressure.

Illustration: H_2 combines with O_2 to form water vapour according to the equation $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$. If 100 mL of hydrogen combine with 50 mL of oxygen, we get 100 mL of water vapour. Thus, the volumes of hydrogen and oxygen which combine together (i.e. 100 mL and 50 mL) bear a simple ratio of 2:1.

5. **Avogadro's Law**: This law was proposed by Amedeo Avogadro. It states that equal volumes of all gases at the same temperature and pressure should contain equal number of moles or molecules.

Illustration: If we take 10L each of NH_3, N_2, O_2 and CO_2 at the same temperature and pressure, all of them contain the same number of moles and molecules.

DALTON'S ATOMIC THEORY

The term atom was first used by John Dalton from the Greek word a-tomio (means indivisible). He proposed the first atomic theory. The **important postulates** of this theory are:

1. Matter is made up of minute and indivisible particles called atoms.
2. Atoms can neither be created nor be destroyed.
3. Atoms of same element are identical in their properties and mass. While atoms of different elements have different properties and mass.
4. Atoms combined to form compound atoms called molecules.
5. When atoms combine, they do so in a fixed ratio by mass.

Dalton's theory could explain the laws of chemical combination.

Atoms and Molecules

Atom is the smallest particle of an element. Molecules are the smallest particle of a substance. A molecule has all the properties of that substance.

Types of molecules

Based on the type of atoms, molecules are divided into two: ***homonuclear molecule and heteronuclear molecule.***

A molecule containing only one type of atom is called *homonuclear molecule*. E.g. H_2, O_2, N_2, O_3 (ozone) etc.

Heteronuclear molecules contain different types of atoms. E.g. CO_2, H_2O, $C_6H_{12}O_6$, NH_3 etc.

Based on the no. of atoms there are three types of molecules: ***monoatomic, diatomic and polyatomic molecules.***

Monoatomic molecules contain only one atom. E.g. all metals, noble gases like He, Ne, Ar etc.

Diatomic molecules contain 2 atoms. E.g. H_2, O_2, N_2, halogens (F_2, Cl_2, Br_2 and I_2)

Polyatomic molecules contain more than two atoms. E.g. ozone (O_3), Phosphorus (P_4), Sulphur (S_8) etc.

Atomic mass

Atomic mass (Relative atomic mass) of an element is a number that expresses how many times the mass of an atom of the element is greater than $1/12^{th}$ the mass of a C^{12} atom.

For e.g. atomic mass of Nitrogen is 14, which means that mass of one N atom is 14 times greater than $1/12^{th}$ the mass of a C^{12} atom.

Atomic mass unit (amu): $1/12^{th}$ the mass of a C^{12} atom is called atomic mass unit (amu).

i.e. 1 amu = $\frac{1}{12}$ x mass of a C^{12} atom

= 1.66×10^{-24} g = 1.66×10^{-27} kg

Nowadays, 'amu' has been replaced by 'u' which is known as ***unified mass***.

Average atomic mass:

Most of the elements have isotopes. So we can calculate an average atomic mass of an element by considering the atomic mass of the isotopes and their relative abundance. For e.g. chlorine has two isotopes ^{35}Cl and ^{37}Cl in the ratio 3:1. So the average atomic mass Cl =(3x35 + 1x37)/4 =35.5

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.

For e.g. molecular mass of H_2SO_4 is calculated as: 2 x 1 + 32 + 4 x 16 = 98 u.

Formula mass:

In the case of ionic compounds (like NaCl), there is no discrete (separate) molecules. Here the positive ions and the negative ions are arranged in a three-dimensional structure. So we can calculate only formula mass by taking molecular formula of the compound.

Mole concept

Mole is the unit of amount of substance. It is defined as the amount of substance that contains as many particles as there are atoms in exactly 12 g C^{12} isotope.

1 mole of any substance contains 6.022 x 10^{23} atoms [602213670000000000000000 atoms]. This number is known as ***Avogadro number or Avogadro constant*** (N_A or N_0).

1 mol of hydrogen atoms = 6.022×10^{23} atoms

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

1 mol of sodium chloride = 6.022×10^{23} formula units of sodium chloride

No. of moles (n) = $\frac{\text{Given mass in gram (w)}}{\text{Molar mass (M)}}$

No. of molecules = no. of moles x 6.022×10^{23}

Molar mass: The mass of one mole of a substance in gram is called its molar mass (gram molecular mass). The molar mass in grams is numerically equal to molecular mass in u.

Molar mass of oxygen = 32g

Molar mass of hydrogen = 2g etc.

Molar volume: It is the volume of 1 mole of any substance. At standard temperature and pressure (STP), molar volume of any gas = 22.4 L (or, 22400 mL). i.e. 22.4 L of any gas at STP contains 1 mole of the gas or 6.022 x 10^{23} molecules of the gas and its mass = molar mass.

For e.g. 22.4 L of hydrogen gas = 1 mole of H_2 = 6.022×10^{23} molecules of hydrogen = 2 g of H_2

Q1) How many moles of water molecules are present in 180 g of water?

Ans: No. of moles $= \frac{\text{Given mass in gram}}{Molar\ mass} = \frac{180}{18} = \underline{\textbf{10 mol}}$

Percentage composition

It is the percentage of each element present in 100g of sample of a substance.

i.e. Percentage composition of an element = $\frac{\text{Mass of that element in the compound x 100}}{\text{Molar mass of the compound}}$

Applications: We can check the purity of a given sample of a substance. Also by knowing the percentage composition, we can calculate the empirical and molecular formula of a compound.

Empirical and Molecular formulae

Empirical formula is the simplest formula of a compound, which gives only the ratio of different elements present in that compound. But molecular formula is the actual formula of the compound, that gives the exact number of different elements present in the sample. For e.g. the empirical formula of glucose is CH_2O but its molecular formula is $C_6H_{12}O_6$.

Molecular formula is related to empirical formula by the equation:

Molecular formula (M.F) = Empirical formula (E.F) x n

Where $n = \frac{\text{Molecular mass (MM)}}{\text{Empirical formula mass (EFM)}}$

By knowing the percentage composition, we can calculate the empirical and molecular formula of a compound as follows:

Q2) An organic compound on analysis gave the following composition. Carbon = 40%, Hydrogen = 6.66% and oxygen = 53.34%. Calculate its molecular formula if its molecular mass is 180.

Ans:

Element	*Percentage*	*Atomic mass*	*Percentage / Atomic mass*	*Simple ratio*	*Simplest whole no. ratio*
C	*40*	*12*	*40/12 = 3.33*	*3.33/3.33 = 1*	*1*
H	*6.66*	*1*	*6.66/1 = 6.66*	*6.66/3.33 = 2*	*2*
O	*53.34*	*16*	*53.34/16 = 3.33*	*3.33/3.33 = 1*	*1*

Empirical Formula = CH_2O

Empirical Formula Mass (EFM) = 12+2+16 = 30

Molar mass (MM) = 180

n = MM/EFM = 180/30 = 6

Molecular formula = Empirical formula x n = $(CH_2O) \times 6 = C_6H_{12}O_6$

Stoichiometry and Stoichiometric calculations

The word 'stoichiometry' is derived from two Greek words – stoicheion (meaning element) and metron (meaning measure). Thus stoichiometry deals with the calculations involving the masses or the volumes of reactants and the products.

Chemical Equation

It is the representation of a chemical reaction by symbols and formulae. Here the reactants are written in the left-hand side and the products, on the right-hand side. (The substances which participate in a chemical reaction are called ***reactants*** and the substances which are formed as a result of a reaction are called ***products***).

A chemical equation should be balanced and the physical states of reactants and products are written in brackets.

The following information are obtained from a chemical equation.

1. An idea about the reactants and products and their physical states.
2. An idea about the masses of reactants and products.
3. An idea about the number moles and molecules of reactants and products.
4. An idea about the volumes of reactants and products at STP.

Limiting reagent (Limiting reactant)

The reagent which limits a reaction or the reagent which is completely consumed in a chemical reaction is called limiting reagent or limiting reactant.

For e.g. in the reaction:

$$2SO_{2(g)} + O_{2(g)} \longrightarrow 2\,SO_{3(g)},$$

2 moles of SO_2 reacts completely with 1 mole of O_2 to form 2 moles of SO_3. If we take 10 moles each of SO_2 and O_2, we get only 10 moles of SO_3 because 10 moles of SO_2 requires only 5 moles of O_2 for the complete reaction. So here SO_2 is the limiting reagent and 5 moles of O_2 remains unreacted.

Q3) A reaction mixture for the production of NH_3 gas contains 250 g of N_2 gas and 50 g of H_2 gas under suitable conditions. Identify the limiting reactant if any and calculate the mass of NH_3 gas produced.

Ans: Nitrogen reacts with Hydrogen to form ammonia according to the equation,

$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$$
$$28g \quad 6g \qquad\qquad 34g$$

28g N_2 requires 6g H_2 for the complete reaction.

So, 250g N_2 requires, 6x 250/28 = 53.57g H_2.

But here there is only 50g H_2.

So we have to consider the reverse case.

i.e. 6g H_2 requires 28g N_2.

So, 50g H_2 requires 28 x 50/6 = 233.33g N_2

Here H_2 is completely consumed. So it is the limiting reagent.

*Amount of ammonia formed = 50+ 233.33 = **283.33 g***

Reactions in solutions

Solutions are homogeneous mixture containing 2 or more components. The component which is present in larger quantity is called solvent and the other components are called solutes. Or, the substance which is dissolved is called solute and the substance in which solute is dissolved is called solvent.
For e.g. in NaCl solution, NaCl is the solute and water is the solvent.

A solution containing only 2 components are called ***binary solution***. If the solvent is water, it is called *aqueous solution*.

The composition of a solution is expressed in terms of concentration. *It is defined as the amount of solute present in a given volume of solution*. Concentration can be expressed in the following ways:

1. **Mass percent (w/w or m/m):** It is defined as the number of parts solute present in 100 parts by mass of solution.

   i.e. Mass % of a component = $\dfrac{\text{Mass of solute } \times 100}{\text{Mass of solution}}$

2. **Mole fraction**: It is defined as the ratio of the number of moles of a particular component to the total number of moles of solution.

   i.e. Mole fraction of a component = $\dfrac{\text{Number of moles of the component}}{\text{Total number of moles of all the components}}$

   For example, in a binary solution, if the number of moles of A and B are n_A and n_B respectively, then mole fraction of the component A (χ_A) = $\dfrac{n_A}{n_A + n_B}$

and mole fraction of the component B (χ_B) = $\dfrac{n_B}{n_A + n_B}$

$$\chi_A + \chi_B = \frac{n_A}{n_A + n_B} + \frac{n_B}{n_A + n_B} = 1$$

i.e. the sum of the mole fractions of all the components in a solution is always equal to 1.
If there are 1,2,3, …….. i components, then $\chi_1 + \chi_2 + \chi_3 + \ldots\ldots\ldots + \chi_i = 1$

3. **Molarity (M):** It is the number of moles of solute dissolved per litre of solution.

i.e. Molarity (M) = $\dfrac{\text{Number of moles of solute (n)}}{\text{Volume of solution in litre (V)}}$

1 M NaOH solution means 1 mole (40 g) of NaOH is present in 1 L of solution.

When a solution is diluted from one concentration to another, its new concentration can be calculated by the equation: $\mathbf{M_1V_1 = M_2V_2}$.

Where M_1 - initial molarity, M_2 - final molarity, V_1 - initial volume and V_2 - final volume.

4. **Molality (m):** It is defined as the number of moles of solute present per kilogram (kg) of the solvent.

i.e. Molality (m) = $\frac{\text{Number of moles of solute}}{\text{Mass of solvent in kg}}$

Among the concentration terms, molarity depends on temperature because it is related to volume, which changes with temperature. All the others are temperature independent.

%%

Q4) Calculate the molarity of a solution containing 8 g of NaOH in 500 mL of water.

Ans: Here mass of solute (NaOH) = 8 g and volume of solution = 500 mL = 0.5 L

Molar mass of NaOH = 40 g mol^{-1}

No. of moles of NaOH = $\frac{\text{mass of NaOH in gram}}{\text{Molar mass of NaOH}} = \frac{8}{40}$ = *0.2 mol*

Molarity = $\frac{\text{Number of moles of solute (n)}}{\text{Volume of solution in litre (V)}} = \frac{0.2}{0.5}$ = ***0.4 M***

Q5) Calculate the mass of oxalic acid dihydrate ($H_2C_2O_4.2H_2O$) required to prepare 0.1M, 250 ml of its aqueous solution.

Ans: Here molarity of solution = 0.1M and volume of solution = 250 mL = 0.25 L

No. of moles of oxalic acid = Molarity x volume of solution in litre = 0.1 x 0.25 = 0.025 mol

Molar mass of Oxalic acid dihydrate = 126 g/mol

Mass of oxalic acid = No. of moles x Molar mass of oxalic acid = 0.025 x 126 = ***3.15 g***

Q6) Calculate the amount of CO_2(g) produced by the reaction of 32g of CH_4 (g) and 32g of O_2 (g).

Ans: $CH_4\ (g) + 2\ O_2\ (g) \longrightarrow CO_2\ (g) + 2\ H_2O\ (g)$

16g 64g 44g 36g

64g O_2 requires 16g CH_4 for the complete reaction.

So, 32g O_2 requires 8g CH_4.

16g CH_4 combines with 64g O_2 to form 44g CO_2.

Therefore, 8g CH_4 combines with 32g Oxygen to form ***22g CO_2***

Q7) If the density of methanol is 0.793 kg L–1, what is its volume needed for making 2.5 L of its 0.25 M solution?

Ans: Density of methanol = 0.793 kg L^{-1}

i.e. mass of 1 L of methanol = 0.793 kg = 793 g

No. of moles of methanol (CH_3OH) = $\frac{Mass\ in\ gram}{Molar\ mass} = \frac{793}{32}$ = *24.78 mol*

Molarity = no. of moles of solute per L of solution = 24.78 M

Here we have to prepare 2.5L (V_2) of 0.25M (M_2) methanol solution from 24.78M (M_1) methanol.

So we can use the equation $M_1V_1 = M_2V_2$

24.78 x V_1 = 0.25 x 2.5 = 0.025 L = 25 mL

Volume of methanol required = ***25 mL***