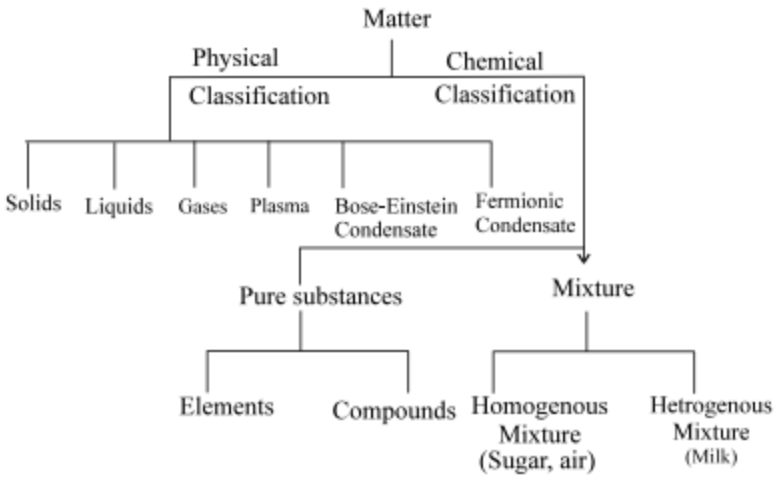
CHAPTER - 01

SOME BASIC CONCEPTS OF CHEMISTRY

SYNOPSIS

1. Classification of matter :



2. Elements: Elements are the purest matter made up of only one kind of atom and can neither be decomposed nor built from simpler substances by any means. The term element was given by Robert Boyle

3. Laws of Chemical Combinations:

(i) **Law of Conservation** of mass was given by French chemist A. Lavoisier (1774) which states that"during any physical or chemical change, the total mass of products is equal to total mass of reactants". It is also called law of indistructubility of matter. It does not hold good for nuclear reactions. He performed careful experimental studies on combustion reactions for reaching the above conclusion

(ii) **Law of definite proportions** was given by Proust (1799) and states that 'A chemical compound always contains same elements combined together in same proportion by mass". For example different samples of CO_2 always have carbon and oxygen in 3:8 ratio by mass. Proust worked with two samples of cupric carbonate

	%Cu	% O	% C
Natural sample	51.35	9.74	38.9
Synthetic sample	51.35	9.74	38.9

(iii) **Law of multiple proportions** was given by John Dalton (1803) and states that "when two elements combine to form two or more compounds the mass of one of the elements which combines fixed mass of the other always form a whole number ratio". This law explains the concept of formation of more than one compound by two elements.

$$H_2 + \frac{1}{2}O_2 \rightarrow H_2O \qquad H_2 + O_2 \rightarrow H_2O_2$$
$$2g \quad 16g \quad 18g \qquad 2g \quad 32g \quad 34g$$

Here the masses of oxygen (16 & 32) which combine with fixed mass of H_2 is 1:2

(iv) **Gay Lussac's law** of Combining Volumes states that when gases react with each other, the volumes bear a simple whole no. ratio to one another and to volume of product (if gases) at similar conditions of pressure and temperature

$$H_2 + \frac{1}{2}O_2 \rightarrow H_2O$$
$$100ml \quad 50ml \quad 100ml$$

(v) **Dalton's Atomic Theory:**

Proposed by John Dalton in 1808. Main points are:

(I) Matter is made of indivisible particles called atoms

(II) Atoms of same elements are identical in physical and chemical properties

(III) Atoms of different elements are different in every respect

(IV) Atoms always combine in whole numbers to form compounds

4. **Drawbacks of Daltons theory :** (a – tomio – indivisible Democritus)

(I) Does not explain structure of atom

(II) Fails to explain binding forces between atoms in compounds

(III) Does not explain Gay Lussac's law

(IV) Does not differentiate between atom and molecule

5. **Avogardo's Law** states that "Equal volumes of all gases, under similar conditions of temperature pressure contain equal number of molecules". Applications are

(I) Deducing atomicity of elementary gases

(II) Deriving relationship between molecular mass and vapour density

(III) Deriving formulas of substance

(IV) Determining molecular wt. of a gas

(V) Deducing the gram molecular wt. of a gas

It was published in the French Journal de physidue

6. **Atom** is the smallest particle of element which may or may not be able to exist independently

7. **Molecule** is the smallest particle of the substance which can exist independently. It can be subdivide as

(1) Homoatomic molecules are molecules of same element and can be further divide monoatomic, diatomic and polyatomic molecules depending on number of atoms. eg. He, O_2 , P_4.

(II) Hetroatomic molecules are molecules of compounds. They can be diatomic and polyatomic

Eg. H_2O, PCl_5, H_2SO_4, NO

8. **Atomic mass unit (a.m.u)** 1 a.m. u or 1U = 1/12 th mass of a C–12 atom (U is unified mass)

9. **Mole** is a unit which represents 6.022×10^{23} particles. The number 6.022×10^{23} is called Avogadro number and is represented by N_A. Avogadro's number of gas molecules occupy a volume of 22400 at

S.T.P. Number of molecules in $1cm^3$ of gas at STP is Loschmidt No. with value 2.688×10^{19}

10. **Atomic Mass**" is the number of times the atom of the element is heavier than H atom" was the proposed definition. Later on Oxygen was preferred as standard. In 1961C – 12 was chosen as standard and thus "the number of times the atom of an element is heavier than $\frac{1}{12}$th part of C–12 is called at mass of the element. Mass spectrometry used to measure atomic masses and molecular mass.

$$\text{Atomic mass} = \frac{\text{Mass of atom of the element}}{\frac{1}{12} \times \text{mass of C} - 12 \text{ atom}}; \quad \text{amu} = \frac{1}{N_A}$$

11. **Average atomic mass** is the mass of each isotope determined separately and then combined in ratio of their occurrence. Suppose a and b are two isotopes of an element with their occurrence ratio p"q

$$\text{then average atomic mass} = \frac{p \times a + q \times b}{p + q}$$

$$\text{eg for C–12, } \frac{0.98892 \times 12 + 0.01108 \times 13.00335 + 2 \times 10^{-12} \times 14.00317}{1} = 12.011 \text{ u}$$

12. Determination of atomic mass:

1. Dulong and Petit's Rule is based on experimental facts."At ordinary temperature, product of atomic mass and specific heat in calories is approximately 6.4 and this product is known as atomic heat of the element".

Atomic mass $\times$ specific heat = 6.4

The law is valid for solid elements except Be, B, Si and C

Correct At. mass = Eq. mass $\times$ valency

(III) Chloride formation method converts the element (whose mass is to be determined) into volatile chloride whose vapour density is found by Victor Mayer method. Molecular mass = 2 $\times$ V.D

(IV) Vapour density method is suitable for elements having volatile chlorides. Atomic mass = Eq. mass of metal $\times$ valency

(V) Mitscherlich's law of isomorphism states that isomorphous substances have similar chemical constitution. Isomorphous substance form crystals of same shape and valencies of elements forming isomorphous salts are also same. Eg. $ZnSO_4$, $7H_2O$ and $FeSO_4$. $7H_2O$ are isomorphous.

13. **Gram Atomic Mass (GAM)** is the mass of an atom expressed in g.

$$\text{No. of gram–atoms of element} = \frac{\text{Mass of element in grams}}{\text{GAM of element}}$$

14. **Molecular mass** is the average relative mass of the molecule as compared with mass of C–12 atom.

$$\text{Molecular mass} = \frac{\text{Average relative mass of one molecule}}{\frac{1}{12} \times \text{Mass of C} - 12 \text{ atom}}$$

$$\text{V.D} = \frac{\text{Mass of V litre of gas}}{\text{Mass of V litre of hydrogen}} \text{ or } \text{V.D} = \frac{1}{2} \frac{\text{Weight of volatile substance}}{\text{Volume at STP}} \times 22400$$

Thus molecules mass = 2 $\times$ V . D

15. **Gram molecular mass or molar mass** is that amount of substance whose mass in grams is equal to molecular mass or the equivalently molecular mass of a substance expressed in grams is called gram molecular mass. Gram molecular mass is also called one gram molecule thus

No. of gam molecules or mole molecules = $\frac{\text{Wt. of substance (in g)}}{\text{GMM of substance}}$

16. **Equivalent Mass** is the number of parts by weight of the substance that combines or displacement, direct or indirectly, 1.008 parts by mass of hydrogen or 8 parts by mass of oxygen or 35.5 parts by mass chlorine.

It can calculated as

(I) Equivalent mass for elements = $\frac{\text{Atomic mass}}{\text{Valency}}$

(II) Equivalent mass for acids = $\frac{\text{Molecular mass}}{\text{Basicity of acid}}$

(III) Equivalent mass for bases = $\frac{\text{Molecular}}{\text{acidity of base}}$

(IV) Equivalent mass for salts = $\frac{\text{Formula mass}}{(\text{Valency of cation})(\text{No.of cations})}$

(V) Equivalent mass for oxidising agents = $\frac{\text{Formula mass}}{\text{No.of electrons gained per molecule}}$

(VI) Equivalent mass for ions = $\frac{\text{Formula mass of radical}}{\text{No.of units of charge}}$

17. **Formula mass** is obtained by adding atomic masses of various atoms present in the formula and term replaces molecular mass in ionic compounds.
18. **Acidity** is the number of OH^- ions that can be displaced from one molecule of a substance.
19. **Basicity** is the number of H^+ ions that can be displaced from one molecule of a substance
20. **Gram equivalent mass (GEM)**

Is the mass of a substance expressed in grams or equivalently the quantity of substance whose mass in grams is equal to its equivalent mass is called one gram equivalent or gram equivalent mass. No. of

gm equivalent = $\frac{\text{Mass in g}}{\text{GEM}}$

PART-I (JEE MAIN)

SECTION-I- Straight objective type questions

1. The total number of atoms of all elements present in one mole of ammonium dichromate is

   1) 19 2) 6.023×10^{23} 3) 114.473×10^{23} 4) 84.322×10^{23}

2. If $\frac{1}{6}^{th}$ mass of a carbon-12 atom is taken to be relative atomic mass unit, the mass of one mole of a substance will

   1) increase by a factor of 2
   2) decrease by a factor of 2
   3) be a function of molecular mass of the substance
   4) remain unchanged.

3. **Assertion :** 22.4 L of N_2 at NTP and 5.6 L of O_2 at NTP contain equal number of molecules.

   Reason : Under similar conditions of temperature and pressure, all gases contain equal number of molecules

   In the light of the above statements, choose the correct option

   1) Both assertion and reason are true and the reason is the correct explanation of the assertion.
   2) Both assertion and reason are true but reason is not the correct explanation of the assertion.
   3) Assertion is true but reason is false.
   4) Assertion and reason both are false.

4. 5 moles of AB_2 weigh $125\times10^{-3}\text{kg}$ and 10 moles of A_2B_2 weigh $300\times10^{-3}\text{kg}$. The molar mass of $A\ (M_A)$ and molar mass of $B(M_B)$ in kg mol^{-1} are:

   1) $M_A=50\times10^{-3}$ and $M_B=25\times10^{-3}$ 2) $M_A=25\times10^{-3}$ and $M_B=50\times10^{-3}$

   3) $M_A=5\times10^{-3}$ and $M_B=10\times10^{-3}$ 4) $M_A=10\times10^{-3}$ and $M_B=5\times10^{-3}$

5. A compound with molar mass 60, has 40% of C, 6.66% of H and remaining oxygen. Calculate molecular formula of the compound.

   1) CH_2O 2) $C_2H_4O_2$ 3) C_3H_8O 4) C_2H_4O

6. In Haber process, 30 litres of dihydrogen and 30 litres of dinitrogen were taken for reaction which yielded only 50% of the expected product. What will be the composition of the gaseous mixture under the aforesaid condition in the end?

   1) 20 litres NH_3, 25 litres N_2, 20 litres H_2

   2) 10 litres NH_3, 25 litres N_2, 15 litres H_2

   3) 20 litres NH_3, 10 litres N_2, 30 litres H_2

   4) 20 litres NH_3, 25 litres N_2, 15 litres H_2

7. A 3M solution of NaCl has density 1.25gmL^{-1}. The molality of the solution is close to
(Given: atomic mass of Na = 23, Cl = 35.5)
1) 2.8 mol/kg 2) 3.8 mol/kg 3) 2.1 mol/kg 4) 3.1 mol/kg

SECTION-II - Numerical Type Questions

8. 10.30 mg of O_2 is dissolved into a liter of sea water of density 1.03 g/mL. The concentration of O_2 in ppm is .....

9. Total mass of CO_2 and H_2O produced from combustion of 2 moles of methane using excess of oxygen is ..... g

10. The moles of a precipitate formed when 3 moles of barium chloride is treated with 3 moles of sodium phosphate is .....

PART-II (JEE ADVANCED)

Section-III - Only one option correct type

11. If a mole was defined to be 3.00×10^{24} (instead of Avogadro's number), what would be the mass of one mole of argon atoms ? (Given: atomic mass of Ar on conventional scale is 40)
A) 199 g B) 199 u C) 40 g D) 40 u

12. Choose correct statement from the following
A) 222 u of $Ca(OH)_2$ contains 32 amu of oxygen
B) If one atom of an element weighs 3.82×10^{-23} g then its atomic mass is 32.68
C) 10 g of $CaCO_3$ contains 0.3 g atom of oxygen
D) Number of atoms in 2 moles of S_8 is greater than that in 55 moles of SO_2

13. How much potassium chlorate must be heated to get as much oxygen as would be obtained from 21.6 g of mercuric oxide ? (Given: atomic mass of Hg = 200, O = 16, K = 39, Cl = 35.5)
A) 3.68 g B) 4.08 g C) 4.76 g D) 5.00 g

14. The most abundant elements in the body of a healthy human adult are oxygen (61.4%), carbon (22.9%), hydrogen (10%) and nitrogen (2.6%). The weight which a 75 kg person would gain if all ^1H atoms are replaced by ^2H atoms is
A) 7.5 kg B) 10 kg C) 15 kg D) 37.5 kg

15. A drop (0.05 mL) of 12 M HCl is spread over a thin sheet of aluminium foil (thickness 1×10^{-2} cm and density $2.7\ gmL^{-1}$). Assuming whole of the HCl is used to dissolve aluminium, the maximum area of hole produced in the foil is
A) $2\times10^{-1}cm^2$ B) $20\times10^{-1}cm^2$ C) $200\times10^{-1}cm^2$ D) $2000\times10^{-1}cm^2$

16. In an experiment 20 mL of a decinormal HCl solution was added to15 mL of a decinormal $AgNO_3$ solution. AgCl was precipitated out and the acid in the solution was neutralised with N/20 NaOH solution. The volume of NaOH required would be :
A) 40 mL B) 20 mL C) 30 mL D) 5 mL

17. A gaseous mixture of propane and butane of volume 3 litre on complete combustion produces 11.0 litre of CO_2 under standard conditions of temperature and pressure. The ratio of volumes of butane to propane in the orginal mixture is:

A) 1:2 B) 2:1 C) 3:2 D) 3:1

Section IV - One or more option correct type

18. Two bulbs A and B contain 16 g of O_2 and 16 g of O_3, respectively, at the same temperature and pressure. Which of the following statements is/are true? (N_A = Avogadro's number)

A) Both bulbs contain same number of atoms

B) Both bulbs contain different number of atoms

C) Both bulbs contain same number of molecules

D) Bulb A contains $\frac{N_A}{2}$ molecules while bulb B contains $\frac{N_A}{3}$ molecules

19. Which of the following statements is/are incorrect ?

A) 6.5 g of a hydrocarbon which contains 7.7% of hydrogen by mass will measure 22.4 L at NTP

B) 18 g of $H_2O(l)$ measures 22.4 L at STP

C) 5.6 L of Cl_2 gas at STP contains 3.011×10^{23} chlorine atoms

D) One molecule of nitric acid contains $5N_A$ atoms

20. A mixture containing 64 g of H_2 and 64 g of O_2 is ignited so that water is formed as follows:

$$2H_2 + O_2 \longrightarrow 2H_2O$$

Identify the correct statement(s) from the following

A) H_2 is the limiting reagent

B) O_2 is the limiting reagent

C) The final mixture contains 72 g of H_2O and 56 g of unreacted H_2

D) The final mixture contains 56 g of H_2O and 72 g of unreacted H_2

21. Two oxides of a metal contain 36.4% and 53.4% of oxygen by mass respectively. If the formula of first oxide is M_2O, then select the correct option(s)

A) Molar mass of metal M is approximately 14 g /mol

B) Formula of the second oxide is M_2O_3

C) Ratio of masses of metal that reacts with a fixed mass of oxygen to form the two oxides is 1:2

D) Formula of the second oxide is MO

22. A mixture of 100 mL of CO, CO_2 and O_2 was sparked. When the resulting gaseous mixture was passed through KOH solution, contraction in volume was found to be 80 mL. The composition of CO, CO_2 and O_2 in the initial mixture may be :

A) 30 mL, 60 mL, 10 mL
B) 30 mL, 50mL, 20 mL
C) 50 mL, 30 mL, 20 mL
D) 20 mL, 70 mL, 10 mL

23. 8 g of oxygen has more number of protons than in:

A) 11 g of CO_2 B) 22 g of CO_2 C) 7 g of CO D) 14 g of CO

Section V - Numerical type questions

24. To a 100 mL $CaCl_2$ solution containing 6.66 g of $CaCl_2$, 100 mL of 0.5M Na_3PO_4 solution is added. The mass of precipitate formed would be —— g

(Given: atomic mass of Ca = 40u, P = 31u, O = 16u)

25. 1120 mL of gaseous HCl at NTP was absorbed in 265 mL of NaOH solution. At the end of reaction, the resulting solution was still alkaline and required 20 mL of 0.15 M H_2SO_4 solution for neutralisation. The normality of NaOH solution is —— N.

26. 1.60 g of a metal were dissolved in HNO_3 to prepare its nitrate. The metal nitrate on strong heating gave 2 g of metal oxide. The equivalent weight of metal is ——

(Given: atomic mass of oxygen is 16 g/mol)

27. Consider the combustion of 1mol of methane using 1 mol of dioxygen. The total number of moles of species present in the final mixture is ——

(Assume that carbon dioxide and water are the only products)

28. What volume of air (in L) at STP containing 21% of oxygen by volume is required to completely burn sulphur (S_8) present in 200 g of a sample that contains 20% of an inert material by wieght?

Sulphur burns according to the reaction : $\frac{1}{8}S_8(S) + O_2(g) \rightarrow SO_2(g)$

(Use molar volume of gas at STP = 22.7L/mol; Atomic mass of S = 32 and O = 16)

Section-VI - Matrix match type

29. Match the following

(M = Molarity, X = Mole fraction, m = Molality, d = Density)

Column - I	**Column - II**
I) 120 g urea in 1L solution $\left(d_{solution} = 1.2g/mL\right)$	P) M = 2
II) 126g HNO_3 + 1.08 lit H_2O $\left(d_{H_2O} = 1g/mL\right)$	Q) $X_{solute} = \frac{1}{31}$
III) 98g H_2SO_4 in 500mL solution $\left(d_{solution} = 1.276g/mL\right)$	R) $X_{solvent} = \frac{30}{31}$
IV) 60g CH_3COOH in 500ml solution $\left(d_{solution} = 1.2g/mL\right)$	S) m = 1.85

A) I - PQRS, II - PQRS, III - PQRS, IV - PQRS

B) I- PS, II - PQRS, III - PS, IV - PS

C) I - PQRS, II - Q,R,S, III - PQRS, IV - PQRS

D) I - PS, II - QRS, III - PS, IV - PS

30. Match the Column - I with Column - II

Column - I	**Column - II**
(I) 0.5 mole SO_2(g)	(P) Occupy 11.2L at NTP
(II) 1g of H_2(g)	(Q) Weighs 24g
(III) 0.5 mol O_3(g)	(R) Number of atoms = $1.5 \times 6.023 \times 10^{23}$
(IV) One mole of O_2(g)	(S) Weighs 32g

A) I - PRS, II - P, III - PQR, IV - S

B) I - RS, II - P, III - QR, IV - S

C) I -PRS, II - R, III - QR, IV - R

D) I - RS, II - R, III - QR, IV - R