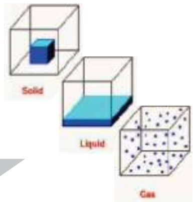

Chapter - 1

Some Basic Concepts Of Chemistry

- **Matter :** Anything that has mass and occupy space.
- **Precision :** If refers to the closeness of various measurements for the same quantity.
- **Accuracy :** It refers to the agreement of a particular value to the true value of the result.
- **Mass and weight :** Mass of a substance is the amount of matter present in it while weight is the force exerted by gravity on an object. The mass of a substance is constant whereas its weight may vary from one place to another due to change in gravity.
- **Volume :** $1 \text{ L} = 1 \text{ dm}^3 = 10^3 \text{cm}^3 = 10^{-3} \text{ m}^3$
- **Temperature :** $K = °C + 273.15; \ \frac{°F - 32}{9} = \frac{°C}{5}$
- **Standard Temperature Pressure (STP) :** 0°C (273.15 K) temperature and 1 atm pressure.
- **Normal Temperature Pressure (NTP) :** 20°C (293.15 K) temperature and 1 atm pressure.
- **Standard Ambient Temperature Pressure (SATP) :** 25°C (298.15 K) temperature and 1 atm pressure
- **Scientific Notation :** Expressing a number in the form $N \times 10^n$, and N can vary b/w 1 to 10.
- **Significant figures :** These are meaningful digits which are known with certainty.
- **Laws of Chemical Combination :**
  - **Law of Conservation of Mass** (Antonie Lavoisier) : Mass can neither be created nor be destroyed.
  - **Law of Definite Proportions** (Joseph Proust) : A given compound

always contains the same elements in the same proportion by mass.

- **Law of Multiple Proportions** (John Dalton) : When two elements combine to form two or more compounds, then the different masses of one element, which combine with a fixed mass of the other, bear a simple ratio to one another.
- **Gay Lussac's Law** : When gases combine or are produced in a chemical reaction, they do so in a simple ratio provided all gases are in the same temperature and pressure.

e.g., $2H_2\,(g) + O_2\,(g) \rightarrow 2H_2O\,(g)$

2 vol 1 vol 2 Vol

(at same T, P)

- **Atomic Mass** : It is defined as the average relative mass of an atom of an element as compared to the mass of an atom of carbon – 12 taken as 12.

  Atomic mass is represented by '*u*' (unified mass).

  $1u = 1.66056 \times 10^{-24}$ g

- **Molecular mass** : It is the sum of the atomic mass of the elements present in the molecule.

  For example : Molecular mass of $CH_4 = (1 \times 12) + (4 \times 1) = 16$ u

- **Avogadro Number** : It is the amount of atoms or molecules present in one mole of a substance.

  Avogadro number $(N_A) = 6.022 \times 10^{23}$

- **Molar Mass :** The mass of one mole of a substance in grams is called its molar mass.

  For example : Molar mass of $CH_4 = (1 \times 12) + (4 \times 1) = 16\text{g mol}^{-1}$

- **Mole (*n*) :** It is amount of a substance that contains as many particles or entities as the number of atoms in exactly 12 grams of pure C-12.

  1 mole of a substance = Molar mass of substance = Avogadro's Number of chemical units = 22.4L volume at STP of gaseous substance

  e.g., 1 mole of CH_4 = 16g of CH_4 = 6.022×10^{23} molecules of CH_4 = 22.4L at STP

  $$n = \frac{wg}{M_m} = \frac{\text{VL (at STP)}}{22.4\text{L}} = \frac{x \text{ particles}}{N_A} = \frac{MV}{1000}$$

- **Molar Volume (V_m) :** It is volume occupied by one mole of any substance.

  Molar volume of a gas = 22.4L at STP (273 K, 1atm) or 22.7L at STP (273

K, 1 bar)

Calculating Molar Volume: PV = nRT

$$\therefore\ V = \frac{nRT}{P} = \frac{1\,mol \times 0.082\,L\ atm\ K^{-1}\ mol^{-1} \times 273\,K}{1\,atm} = 22.4\,L$$

Or

$$V = \frac{nRT}{P} = \frac{1\,mol \times 0.083\,L\ bar\ K^{-1}\ mol^{-1} \times 273\,K}{1\,atm} = 22.7\,L$$

- **Percentage Composition :** Mass % of the element

$$= \frac{\text{Mass of element in a molecule of the compound} \times 100}{\text{Molecular mass of the compound}}$$

- **Empirical Formula :** It represents the simplest whole number ratio of various atoms present in a compound. *e.g.*, CH is the empirical formula of benzene.
- **Molecular Formula** : It shows the exact number of different of atoms present in a molecule of a compound. *e.g.*, C_6H_6 is the molecular formula of benzene.
- **Relationship between empirical and molecular formulae :** Molecular formula = n × Empirical formula

Where; $n = \dfrac{\text{Molar mass}}{\text{Empirical formula mass}}$

- **Information Conveyed by a chemical equation :**

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

(i) 1 molecule of N_2 +3 molecules of H_2 → 2 molegules of NH_3

(ii) 1 mole of N_2 + 3 mole of H_2 → 2 mole of NH_3

(iii) 1 × 28g of N_2 + 3 × 2 g of H_2 → 2 × 17 g of NH_3

(iv) 1 × 22.4L of N_2 at STP + 3 × 22.4L of H_2 at STP → 2 × 22.4L of NH_3 at STP

- **Limiting Reagent** : It is the reactant which gets consumed first or limits the amount of product formed.
- **Mass Percent** : It is the mass of the solute in grams per 100 grams of the solution.

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

- **Parts per million (ppm) :** It is part of solute per million part of solution by mass.

$$\text{ppm} = \frac{\text{Parts of solute (by mass)} \times 10^6}{\text{Parts of solution (by mass)}}$$

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

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

Molarity equation : $M_1V_1 = M_2V_2$

(Before dilution) (After Dilution)

Malarity of a solution decreases on increasing temperature.

Malarity of pure water is 55.56 mol L^{-1}

- **Molality (m)**—It is number of moles of solute dissolved per 1000g (1kg) of solvent.

$$\text{Molality} = \frac{\text{No. of moles of solute}}{\text{mass of solution in kg}}$$

Molality is independent of temperature.

- **Mole Fraction(x)** is the ratio of number of moles of one component to the total number of moles (solute and solvents) present in the solution.

$$x_1 = \frac{n_1}{n_1 + n_2} \quad \text{and } x_2 = \frac{n_2}{n_1 + n_2}$$

The sum of all the mole fractions in a solution is equal to one. *i.e.*, $x_1 + x_2 = 1$

Importance of Chemistry & Nature of Matter

1-Mark Questions

1. Name two chemical compounds used in treatment of cancer.
2. What is AZT ? Write its use.
3. Give an example each of homogeneous and heterogeneous mixture.
4. Differentiate solids, liquids & gases in terms of volume & shapes.
5. Classify following as pure substances and mixtures : Air, glucose, gold,

sodium and milk.

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

Properties of matter and their Measurement

7. What is the SI unit of density ?
8. What is the SI unit of molarity ?
9. Define accuracy and precision.
10. What are the two different system of measurement ?
11. What is the difference between mass & weight ?

Uncertanity in Measurement

12. Define significant figures.
13. Define accuracy and precision
14. Which measurement is more precise 4.0g or 4.00g ? [Ans. 4.00 g]
15. How many significant figures are there in (i) 3.070 and (ii) 0.0025 ? [**Ans.** (i) 4 (ii) 2]
16. Express the following in the scientific notation : (i) 0.0048 (ii) 234,000

Laws of Chemical Combinations & Dalton's Atomic Theory

17. State Avogadro's law.
18. State law of definite proportions.
19. State Gay Lussac's Law of combining volumes of gases.
20. If ten volumes of dihydrogen gas react with five volumes of dioxygen gas, how much volume of water vapour would be produced ? [**Ans.** 10 volumes]

Atomic and Molecular masses and Mole Concept

21. Define unified mass (u).
22. Calculate the number of atoms in 32.0 *u* of He. [**Ans.** 8]
23. Define molar volume of a gas.
24. What is the volume of 17 g of NH_3 gas at STP (298 K, 1 atm) ? [**Ans.** 22.4 L]
25. What is the value of one mole ?

26. Calculate the number of molecules present in 22.0 g of CO_2.

[**Ans.** 3.011×10^{23}]

27. How many molecules of SO_2 are present in 11.2 L at STP ?

[**Ans.** 3.011×10^{23}]

28. Which has more number of atoms ? 1.0 g Na or 1.0 g Mg. [**Ans.** 1.0 g Na]

29. How many oxygen atoms are present in 16 g of ozone (O_3) ?

[**Ans.** 2.007×10^{23}]

30. At STP, what will be the volume of 6.022×10^{23} molecules of H_2 ?

[**Ans.** 22.4L]

31. 1L of a gas at STP weighs 1.97g. What is molecular mass ?

[**Ans.** 44.128 g mol^{-1}]

Percentage Composition, Empirical and Molecular Formula

32. Write the relationship between empirical formula and molecular formula.

33. Which is more informative ? Empirical formula or Molecular formula.

34. A subtance has molecular formula $C_6H_{12}O_{6.}$ What is its empirical formula.

35. Empirical formula of a compound X(Molar mass = 78 mol^{-1}) CH. Write its molecular formula.

Stochiometry and Stoichiometric Calculations

36. How are 0.5 mol Na_2CO_3 and 0.5 M Na_2CO_3 different from each other ?

37. Why molality is preferred over molarity of a solution ?

38. Define molarity of a solution.

39. What is the effect of temperature on molarrity of solution ?

40. What is limiting reactant in a reaction ?

Importance of Chemistry & Nature of Matter

2 Mark Questions

1. How can we say that sugar is solid and water is liquid?
2. How is matter classified at macroscopic level?
3. Classify following substances as element, compounds and mixtures : water, tea, silver, steel, carbon dioxide and platinum.

Properties of matter and their Measurement

4. The body temperature of a normal healthy person is 37°C. Calculate its value in°F.

5. At what temperature will both the Celsius and Fahrenheit scales read the same value?

6. Convert 5L into m^3.

7. What does the following prefixes stand for : (a) pico (b) nano (c) micro (d) deci

Uncertanity in Measurement

8. How many significant figures are present in the answer of the following calculations :

   (i) $0.0125 + 0.8250 + 0.025$ (ii) $\dfrac{0.025 \times 298.15 \times .1155}{0.5785}$

9. Convert '450 pm' into SI unit and write the answer in scientific notation upto 2 significant figures.

   [**Ans**. 4.5×10^{-10} m]

10. The density of vanadium is 5.96 g cm^{-3}. Express this in SI unit.

    [**Ans.**5960 kg m^{-3}]

Laws of Chemical Combinations & Dalton's Atomic Theory

11. 45.4 L of dinitrogen reacted with 22.7 L of dioxygen and 45.4 L of nitrous oxide was formed. The reaction is given below : $2\,N_2\,(g) + O_2\,(g) \rightarrow 2\,N_2O\,(g)$ Which law is being obeyed in this experiment? Write the statement of the law.

12. Write main points of Dalton's Atomic Theory.

Atomic and Molecular masses and Mole Concept

13. Give one example each of a molecule in which empirical formula and molecular formula is (i) Same (ii) Different.

14. Calculate the number of moles in the following masses :

    (i) 7.85g of Fe; (ii) 7.9mg of Ca

15. Calculate average atomic mass of chlorine using following data:

Isotope	% Natural abundance	Molar mass
^{35}Cl	75.77	34.9689
^{37}Cl	24.33	36.9659

[**Ans.** 35.5 u]

Percentage composition, empirical and molecular formula

16. Give one example of molecule in whch empirical formula and molecular formulaare (i) same (ii) different.

17. Calculate the present of carbon, hydrogen and oxygen in ethanol (C_2H_5OH)

[**Ans.** 52.14%, 13.13%, 34.73%]

18. How much copper can be obtained from 100 g of $CuSO_4$?[**Ans.** 39.8g]

Stiochiometry and Stoichiometric Calculations

19. Calculate the amount of water (*g*) produced by the combustion of 16 g of methane. [**Ans.** 36g]

20. How many moles of methane are required to produce 22 g CO_2 (g) after combustion? [**Ans.** 0.5 mol]

21. A solution is prepared by adding 2 g of a substance A to 18 g of water. Calculate the mass per cent of the solute. [**Ans.** 10%]

22. Calculate molarity of water if its density is 1.00 g mL^{-1}. [**Ans.** 55.56 M]

23. Calculate the molarity of NaOH in the solution prepared by dissolving its 4 g in enough water to form 250 mL of the solution. [**Ans.** 0.4 M]

24. The density of 3 M solution of NaCl is 1.25 g mL^{-1}. Calculate molality of the solution. [**Ans.** 2.8m]

25. Calculate the molarity of a solution of ethanol in water in which the mole fraction of ethanol is 0.040 (assume the density of water to be one).

[**Ans.** 2.31 M]

26. NH_3 gas can be prepared by Haber's process as, $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$. At a particular moment concentration of all the species is 2 moles; calculate the concentration of N_2 and H_2 taken initially.

[**Ans.** 3 mole, 5 moles]

27. A sample of drinking water was found to be severely contaminated with

chloroform, $CHCl_3$, supposed to be carcinogenic in nature. The level of contamination was 15 ppm (by mass).

(i) Express this in percent by mass.

(ii) Determine the molality of chloroform in the water sample.

[**Ans.** (i) ~ 15×10^{-14} g (ii) 1.25×10^{-4} m]

28. Potassium superoxide, KO_2 is used in rebreathing gas masks to generate oxygen.

$$4KO_2(s) + 2H_2O(l) \rightarrow 4KOH(s) + 3O_2(g)$$

If a reaction vessel contains 0.15 mol KO_2 and 0.10 mol H_2O, what is the limiting reactant? How many moles of oxygen can be produced?

[**Ans.** KO_2, 1.125 mol]

29. How many grams of HCl react with 5.0 g of MnO_2 according to the equation.

$4HCl(aq) + MnO_2(s) \rightarrow 2H_2O(l) + MnCl_2(aq) + Cl_2(g)$ [**Ans.** 8.40 g]

30. 0.5 mol of H_2S and SO_2 are mixed together in a reaction flask in which the following reaction takes place : $2H_2S(g) + SO_2(g) \rightarrow 2H_2O(l) + 3S(s)$

Calculate the number of moles of sulphur formed. [Ans. 0.75 mol]

31. Pure oxygen is prepared by thermal decomposition of $KC1O_3$ according to the equation :

$$KClO_3(s) \xrightarrow{\Delta} KCl(s) + \frac{3}{2}O_2(g)$$

Calculate the volume of oxygen gas librated at STP by heating 12.25 *g* $KClO_3$(s). [**Ans.** 3.36 L]

Importance of Chemistry and Natureof Matter

3-Marks Questions

1. Give three main points of difference between a compound and a mixture.
2. Define homogeneous and heterogeneous mixture with example.

Properties of matter and their Measurement

3. Write seven fundamental quantities & their units
4. Pressure is defined as force per unit area of the surface. The SI unit of

pressure, Pascal is :

1Pa $= 1\ Nm^{-2}$

If mass of air at sea level is 1034 g cm^{-2}, calculate the pressure in Pascal.

[**Ans.** 1.01332×10^5 Pa]

Laws of Chemical Combinations & Dalton's Atomic Theory

5. The following data are obtained when dinitrogen and dioxygen react together to form different compounds :

	(i)	**(ii)**	**(iii)**	**(iv)**
Mass of dinitrogen	14	14	28	28
Mass of dioxygen	16	32	32	80

Which law of chemical combination is obeyed by the above experimental data ? Give its statement.

Atomic and Molecular Masses and Mole Concept

6. Calculate :

(i) Mass in gram of 5.8 mol N_2O

(ii) No of moles in 8.0 g of O_2

(iii) Molar mass if 11.2 L at STP weigh 8.5 g.

[**Ans.** (i) 255.2 g (ii) 0.25 mol (iii) 17 g mol^{-1}]

7. In three moles of ethane (C_2H_6), calculate the following :

(i) No of moles of carbon atom,

(ii) No of moles of hydrogen atoms,

(iii) No of molecules of ethane.

[**Ans.** (i) 6 moles, (ii) 18 moles, (iii) 1.81×10^{24}]

8. 16 g of an ideal gas SO_x occupies 5.6 L at STP. What is its molecular mass? What is the value of X ? [**Ans.** 64u, $x = 2$]

9. Calculate the number of moles :

(i) 5.0 L of 0.75 M Na_2CO_3

(ii) 7.85 g of Fe

(iii) 34.2 of sucrose ($C_{12}H_{22}O_{11}$) [**Ans.** (i) 3.75, (ii) 0.14, (iii) 0.1]

10. Calculate the number of atoms in each of the following :

(i) 52 moles of Ar. (ii) 52u of He (iii) 52g of He.

[**Ans.** (i) 3.13×10^{25} (ii) 13 (iii) 7.83×10^{24}]

Percentage Composition, Empirical and Molecular Formula

11. Vitamin C is essential for the prevention of scurvy. Combustion of 0.2000g of vitamin C gives 0.2998g of CO_2 and 0.819g of H_2O. What is the empirical formula of vitamin C ?

12. 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 formulas? [**Ans.** CH_2C1, $C_2H_4Cl_2$]

13. A compound made up of two elements A and B has A = 70%, B = 30%. Their relative number of moles in the compound is 1.25 and 1.88, calculate :

(i) Atomic masses of the elements A and B

(ii) Molecular formula of the compound , if its molecular mass is found to be 160. [**Ans.** (i) 56 and 16, (ii) A_2B_3]

Stoichiometry and Stoichiometric Calculations

14. Calculate the mass of sodium acetate (CH_3COONa) required making 500 mL of 0.375 molar aqueous solution. (Molar mass of sodium acetate is 82.0245 g mol^{-1}). [**Ans.** 15.375 g]

15. Calculate the concentration of nitric acid in moles per litre in a sample which has a density, 1.41 g mL^{-1} and the mass per cent of nitric acid in it being 69%. [**Ans.** 15.44 M]

16. What is the concentration of sugar ($C_{12}H_{22}O_{11}$) in mol L^{-1} if its 20 g are dissolved in enough water to make a final volume up to 2L ? [**Ans.** 0.029 M]

17. Calcium carbonate reacts with aqueous HC1 according to the reaction :

$CaCO_3\ (s) + 2\ HC1\ (aq) \rightarrow CaCl_2\ (aq) + CO_2\ (g) + H_2O(l)$

What mass of $CaCO_3$ is required to react completely with 25 mL of 0.75 M HC1 ? [**Ans.**0.94 g]

18. The reaction $2C + O_2 \rightarrow 2CO$ is carried out by taking 24.0 g of carbon and 96.0 g of O_2. Find out.

(i) Which reactant is left in excess ?

(ii) How much of it is left ?

(iii) How many grams of the other reactant should be taken so that nothing is left at the end of the reaction ? [**Ans.** (i) O_2, (ii) 64 g, (iii) 72]

HOTS Question

19. A 10 g sample of a mixture of calcium chloride and sodium chloride is treated with Na_2CO_3 to precipitate calcium as calcium carbonate. This $CaCO_3$ is heated to convert all the calcium to CaO and the final mass of

CaO is 1.62 g. Calculate % by mass of NaCl in original solution.

[**Ans.** 67.9%]

Atomic and Molecular masses and Mole Concept

5-Mark Questions

1. (i) A black dot used as a full stop at the end of a sentence has a mass of about one attogram. Assuming that the dot is made up of carbon, calculate the approximate number of carbon atoms present in the dot.

*[**Hint :** 1 attogram = 10^{-18} g]* [**Ans.** 5.02×10^4]

(ii) Which one of the following will have largest number of atoms ?

(a) 1g Au (s) (b) 1g Na (s) (c) 1g Li (s) (d) 1g of Cl_2(g)

[**Ans.** (i) 39.81 g (ii) 1 g of Li]

Percentage Composition, Empirical and Molecular Formula

2. (i) What is the difference between empirical formula and molecular formula ?

(ii) A welding fuel gas contains carbon and hydrogen only. Burning a small sample of it in oxygen gies 3.38 g carbon dioxide 0.690 g of water and no other products. A volume of 10.0 L (measured at STP) of this welding gas is found to weigh 11.6 g. Calcuate (i) empirical formula, (ii) molar mass of the gas, and (iii) molecular formula.

[**Ans.** (i) CH, (ii) 26 g mol^{-1}, (iii) C_2H_2]

Stoichiometry and Stoichiometric Calculations

3. (i) What is the difference between Molarity and Molality.

(ii) The Molarity of a solution of sulphuric acid is 1.35 M. Calculate its molality. (The density of acid solution is 1.02 g cm^{-3}).[**Ans.** 1.52 m]

4. (i) Define : (a) Mole fraction (b) Mass percentage.

(ii) 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.** 0.0025 L]