SECTION - I : STANDARD - XI

CHAPTER

1

SOME BASIC CONCEPTS OF CHEMISTRY

POINTS TO REMEMBER

1. Introduction

- Chemistry is the branch of science. It plays an important role to fulfil our basic need. i.e. food, clothes and shelter.
- The subject chemistry is the study of composition, structure and properties of matter and the reactions by which one form of matter may be converted into another form.
- Chemistry is mainly divided into five branches as : (1) Physical chemistry (2) Inorganic chemistry (3) Organic chemistry (4) Analytical chemistry (5) Biochemistry

2. Importance and Scope of Chemistry

- Chemistry is an important branch of science. It has wide scope in different fields. We apply chemistry in day to day life. Knowingly or unknowingly people use chemistry in kitchen. It has wide scope in industries like engineering, pharmaceutical, food industry, textile industry, laundry, agriculture, perfumery, cosmetics, polymer, printing, paper industry, etc.

3. Historical Approach to Particulate Nature of Matter

- **Units :** The value of a physical quantity is expressed as its mathematical value and a unit/units are expressed in **CGS** (centimeter, gram and second), **FPS** (foot, pound, second), **MKS** (meter, kilogram, second), **SI** (System International).
- The SI system has **seven fundamental units.**

Quantity	SI unit	Symbol
1. Length	meter	m
2. Mass	kilogram	kg
3. Time	second	s
4. Temperature	kelvin	K
5. Amount of substance	mole	mol
6. Electric Current	ampere	A
7. Luminous Intensity	candela	cd

- **Derived unit :** The units of all physical quantities can be derived from the seven fundamental units are known as derived unit.

(ii) **Law of constant composition :** This law states that *a pure chemical compound always contains same elements combined together in the same proportion by weight.*

(iii) **Law of multiple proportion :** According to this law, *when two elements combine to form two or more than two compounds, then the weights of one of these elements which combine with the fixed weight of other element bear a simple ratio to one another.*

(iv) **Law of reciprocal proportion :** This law states that *weights of two elements A and B, which separately combine with a third element C, are either the same or simple multiple of the weights of A and B which combine with each other.* This is also called law of equivalent proportions.

(v) **Gay Lussac's law of combining volumes :** This law states that *under similar conditions of temperature and pressure, whenever gases react together, the volumes of the reacting gases as well as products (if gases) bear a simple whole number ratio.*

5. Dalton's Atomic Theory

The main postulates of Dalton's atomic theory are :

(i) All matter is made up of very tiny particles which cannot be further broken. These tiny particles are called atoms.

(ii) Atoms can neither be created nor destroyed during chemical reactions.

(iii) Atoms of a particular element are identical in all respects, i.e., have same size, mass and other properties.

(iv) Atoms of different elements are different.

(v) Atoms combine to form molecules.

6. Concept of Elements, Atoms and Molecules

Atoms and Molecules : Scientists quest to know the ultimate particle of a substance has led to the discovery of atoms and molecules.

Atom : The smallest particle of an element which may or may not have independent existence is called an atom.

Every atom of an element has a definite mass of the order of 10^{-26} kg and has spherical shape of radius of the order of 10^{-15} m.

Mass of hydrogen is 1.667×10^{-26} kg.

Atoms may or may not have free existence.

Molecule : The smallest particle of a substance (element or compound) which is capable of independent existence is called a molecule.

7. Atomic and Molecular Masses

Atomic mass : Atoms are extremely small particles and it is very difficult to determine their actual masses. For example, mass of one atom of hydrogen has been found to be = 0.00000000000000000000000166 g or 1.66×10^{-24} g. Obviously, such a small mass cannot be determined even with the help of a very sensitive balance. Moreover, this value is not very convenient for calculations. To overcome this problem, the atomic masses are expressed as relative masses, i.e. with reference to mass of a standard atom.

The stable isotope of carbon (C^{12}) has been selected as the standard for comparing the atomic and molecular masses of elements and compounds. The atomic mass of the standard, the isotope ^{12}C of carbon is chosen to be 12. Thus, atomic mass may be defined as :

The relative mass of an atom of the element as compared to an atom of carbon (C^{12}) taken as 12.

In other words, atomic mass expresses as to how many times an atom of the element is heavier than a carbon atom taken as 12 or $\frac{1}{12}$th of the mass of a carbon atom. Therefore,

$$\text{Atomic mass} = \frac{\text{Mass of an atom}}{\frac{1}{12}\text{th mass of a carbon atom } (C^{12})}$$

This scale of relative masses of atoms is called *atomic mass unit scale* and is abbreviated as *a.m.u.* The quantity of mass equal to $\frac{1}{12}$ of the mass of an atom of carbon (C^{12}) is *atomic mass unit*. Recently a.m.u. is replaced by 'u' meaning united mass.

Gram atomic mass : *The quantity of an element whose mass in grams is numerically equal to its atomic mass* is called **gram atomic mass** or **gram atom.** In simple words, atomic mass of an element expressed in grams is the gram atomic mass.

e.g. Gram atomic mass of oxygen = 16 g

Average atomic mass : The observed atomic mass of the atom of an element is the average atomic mass of the elements taking into consideration the natural abundance of the isotope.

Isotops are the atoms of the same element having same atomic number but different mass number.

Molar mass : Like atomic masses, molar masses are also expressed relative to the stable isotope of carbon (C^{12}) having mass number 12. Thus, molar mass may be defined as *the average relative mass of its molecule as compared to the mass of an atom of carbon (C^{12}) taken as 12.*

Physical Quantity	Relation with Other Basic Quantities	SI Unit
1. Area	Length Square	m^2
2. Volume	Length Cube	m^3
3. Density	Mass per unit Volume	kg/m^3 OR $kg\,m^{-3}$
4. Velocity	Distance travelled per unit time	m/s OR ms^{-1}
5. Acceleration	Velocity change per unit time	$m/s \times \frac{1}{s}$ OR ms^{-2}
6. Force	Mass × Acceleration	$kg\,m/s^2$ OR $kg\,ms^{-2}$ (= newton, N)
7. Pressure	Force per unit area	$kg\,m^{-1}s^{-2}$ (= pascal, Pa)
8. Energy	Force × Distance travelled	$kg\,ms^{-2} \times m$ OR $kg\,m^2s^{-2}$ (= joule, J)
9. Concentration	Mole per cubic metre	$mol\,m^{-3}$
10. Electric Charge	Current × Time	A·S (= coulomb, C)
11. Electric potential or potential difference	Energy per unit charge	$J\,A^{-1}s^{-1}$ (= volt, V) OR $J\,C^{-1}$ OR $kg\,m^2\,s^{-3}\,A^{-1}$
12. Electrochemical Equivalent	Z = E/F	kg/C OR $kg\,C^{-1}$
13. Heat Capacity	$C_P = dH/dT$ OR $C_V = dE/dT$	$J\,K^{-1}\,mol^{-1}$

- Some common SI prefixes.

Prefix	Symbol	Magnitude in 'm'
1. Tera-	T	10^{12}
2. Giga-	G	10^9
3. Mega-	M	10^6
4. myria-	my	10^4
5. kilo	k	10^3
6. hect-	h	10^2
7. deca-	da	10
8. deci-	d	10^{-1}
9. centi-	c	10^{-2}
10. milli-	m	10^{-3}
11. micro-	μ	10^{-6}
12. nano-	n	10^{-9}
13. pico-	p	10^{-12}
14. femto-	f	10^{-15}

- **Matter :** All the objects around us are composed of matter. Matter is defined as anything that occupies space and has mass.
- **Classification of Matter :**

Physical classification - solid, liquid and gaseous state.

Chemical classification - elements, compounds and mixtures.

- **Element :** Element is the simplest pure form of matter. It neither decompose by simple chemical means nor formed by the combination of other substances. Till date 118 elements are identified, 92 elements are present in nature and about 26 elements are synthesised in laboratory.
- **Compound :** A substance made up of two or more elements chemically combined in a fixed proportion by weight.
- **Mixture :** A material obtained by mixing two or more substances (elements or compounds) in any proportion is called mixture.

If two or more phases are present in a mixture then it is called heterogeneous mixture.

If all the constituents are present in one phase and the concentration of the constituent remain uniform throughout the mixture then it is called homogeneous mixture.

4. Laws of Chemical Combination

(i) **Law of conservation of mass :** This law states that *during any physical or chemical change, the total mass of the products is equal to the total mass of the reactants.* This law is also known as law of indestructibility of matter.

$$\text{Molecular mass} = \frac{\text{Mass of a molecule}}{\frac{1}{12}\text{th mass of a carbon atom } (C^{12})}$$

Gram molecular mass : *The quantity of substance whose mass in grams is numerically equal to its molecular mass* is called **gram molecular mass.** In other words, molecular mass of a substance expressed in grams is called gram molecular mass. It is also called **gram molecule.**

e.g. Gram molecular mass of CO_2 = 44 g

8. Avogadro's Law, Avogadro's Number and Mole Concept

Avogadro's Law : Equal volumes of all gases under identical conditions of temperature and pressure contain equal number of molecules.

OR

Avogadro's law is stated as, at constant temperature and pressure, volume of the gas is directly proportional to the number of molecules.

At constant P and T,

$$V \propto n$$

$$\text{i.e. } \frac{V}{n} = \text{Constant}$$

According to general gas equation,

$$PV = nRT$$

$$\therefore \frac{V}{n} = \frac{RT}{P} \text{ at STP}$$

$T = 273.15$ K, $P = 1$ atm

$R = 0.082051$ atm mol^{-1} K^{-1}

$$\therefore \frac{V}{n} = \frac{0.082051 \times 273.15}{1} = 22.414 \text{ L mol}^{-1}$$

This value is called Avogadro's molar volume or molar gas volume at STP.

- **Mole Concept**

  A **mole** *is a collection of* 6.022×10^{23} particles. It may also be defined as the amount of substance which contains the same number of elementary units (atoms, molecules or ions) as the number of atoms present in 12 g of carbon (C^{12}).

  Mole and Gram atomic mass : Mass of one mole of atoms (6.022×10^{23} atoms) of any element in grams is equal to its gram atomic mass.

  One mole atoms = 6.022×10^{23} atoms

  = Gram atomic mass of element

  Mole and Gram molecular mass : Mass of one mole (mass of 6.022×10^{23} molecules) of a substance in grams is equal to its gram molecular mass.

  One mole molecules = 6.022×10^{23} molecules

  = Gram molecular mass

  In case of ionic compounds, the mass of 1 mole of formula units (6.022×10^{23} formula units) of any ionic substance in grams is equal to its gram formula unit.

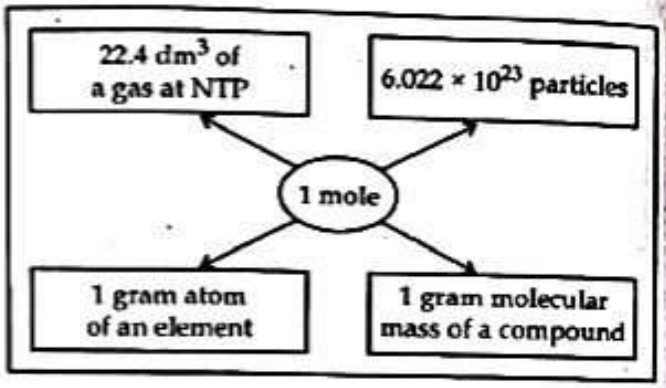
  Mole in terms of volume : The volume occupied by one mole of a gas at N.T.P. is known as gram molar volume. It is equal to 22.4 litres or 22.4 dm^3.

  e.g. 1 mole of O_2 gas at N.T.P. occupies

  $= 22.4 \text{ dm}^3 = 0.0224 \text{ m}^3$

  1 mole of SO_2 gas at N.T.P. occupies

  $= 22.4 \text{ dm}^3 = 0.0224 \text{ m}^3$



- (i) $\text{Number of moles (n)} = \frac{\text{Mass of the substance}}{\text{Molar mass of the substance}}$

  (ii) Number of molecules

  = Number of moles × Avogadro number

  $= n \times 6.022 \times 10^{23}$

  (iii) Volume of n moles of the gas at N.T.P.

  = Number of moles × 22.4 dm^3 = $n \times 22.4$ dm^3

9. Percentage Composition and Empirical and Molecular Formula

- For characterization and identification of any compound, determination of percentage composition of constituent atoms of an element is essential.

$$\text{Percentage of an element} = \frac{\text{Mass of the element in 1 mole of the compound}}{\text{Molar mass of the compound}} \times 100$$

- From the percentage composition empirical and molecular formula can be determined.

- Empirical formula of a compound is defined as a chemical formula indicating relative number of constituent atoms in the simplest ratio.
- The molecular formula indicates the actual number of constituent atoms in a molecule. **Molecular formula is integral multiple of empirical formula.**
- **Steps for determination of empirical and molecular formula of a compound :**

(i) Percentage of all the elements present in the compounds are determined experimentally. The sum of the percentages of the constituent elements must be hundred. If it is less than 100% then oxygen is present with percentage equal to the difference between 100 and the sum of percentage of constituent elements.

(ii) $$\text{Number of moles of the element} = \frac{\text{Percentage of the element}}{\text{Atomic mass of the element}}$$

(iii) The ratio of number of moles of each constituent element is determined and the ratio is converted into smallest simple whole number.

e.g. In case of acetic acid (CH_3COOH) the smallest simple whole number ratio is C : H : O is 2 : 4 : 2.

(iv) If the ratio obtained is fractional, then it is multiplied by convenient integer to obtain whole number ratio.

(v) The elements are written side by side and then the number of moles of elements are written as the subscripts to the right side of the element.

(vi) Knowing molecular mass of the compound 'n' can be calculated

$$n = \frac{\text{Molecular mass of the compound}}{\text{Empirical mass of the compound}}$$

(vii) Molecular formula of the compound

= n × empirical formula of the compound

10. Chemical Reactions, Stoichiometry and Calculation Based on Stoichiometry

- **Chemical reaction** is a process in which one or more substances interact with each other to produce one or more substances.
- **Stoichiometry** is a process of making calculation based on formulae and balanced chemical equation.

e.g. Nitrogen + Hydrogen → Ammonia

$$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$$

3 and 2 are coefficients of hydrogen and ammonia are called stoichiometric coefficient.

- Balance chemical reaction can be written in 3 steps.

Step I : Write the chemical reaction in word.

e.g. Methane + Oxygen → Carbon dioxide + Water

Step II : Write the chemical equation in terms of chemical formula of each substance (reactant and product).

$$CH_{4(g)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(g)}$$

Step III : Balance the mass of chemical reaction by using proper whole number coefficient for each reactant and product.

$$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)}$$

- **Limiting and Excess reactants :** Generally, as a cost saving measure the cheaper reactant is taken in excess and other reactants are used in lesser amount. The reactant that is taken in lesser amount get consumed earlier and the reaction stops; this reactant is called as **limiting reactant.** The excess reactant which is present in excess is called excess reactant.