

## UNIT 1

### SOME BASIC CONCEPTS OF CHEMISTRY

**Chemistry:** Chemistry is the branch of science that deals with the composition, structure and properties of matter. Chemistry is called the science of atoms and molecule

#### 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.

**PROPERTIES OF MATTER AND THEIR MEASUREMENT**--Every substance has unique or characteristic properties. These properties can be classified into two categories – **physical properties** and **chemical properties**.

**Physical properties** are those properties which can be measured or observed without changing the identity or the composition of the substance. E.g. colour, odour, melting point, boiling point, density etc.

The measurement or observation of **chemical properties** requires a chemical change to occur. e.g. Burning of Mg-ribbon in air

**Chemical properties** are characteristic reactions of different substances; these include acidity or basicity, combustibility etc. Many properties of matter such as length, area, volume, etc., are quantitative in nature.

**Metric System** was based on the decimal system.

#### **The International System of Units (SI)**

The International System of Units (in French Le Systeme International d'Unites--abbreviated as SI) was established by the 11th General Conference on Weights and Measures (CGPM from *Conference Generale des Poids at Measures*). The SI system has seven *base units*

| Quantity            | Unit     | Symbol |
|---------------------|----------|--------|
| Length              | metre    | m      |
| Mass                | kilogram | kg     |
| Time                | second   | s      |
| Temperature         | kelvin   | K      |
| Amount of substance | mole     | mol    |
| Electric current    | ampere   | A      |
| Luminous intensity  | candela  | Cd     |

## Prefixes in SI system

| Multiple   | Prefix | Symbol |
|------------|--------|--------|
| $10^{-12}$ | pico   | p      |
| $10^{-9}$  | nano   | n      |
| $10^{-6}$  | micro  | $\mu$  |
| $10^{-3}$  | milli  | m      |
| $10^{-2}$  | centi  | c      |
| $10^{-1}$  | deci   | d      |
| 10         | deca   | da     |
| $10^2$     | hecto  | h      |
| $10^3$     | kilo   | k      |
| $10^6$     | mega   | M      |
| $10^9$     | giga   | G      |
| $10^{12}$  | tera   | T      |

**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. The mass of a substance can be determined very accurately by using an analytical balance

**Volume--** Volume has the units of  $(\text{length})^3$ . So volume has units of  $\text{m}^3$  or  $\text{cm}^3$  or  $\text{dm}^3$ . A common unit, litre (L) is not an SI unit, is used for measurement of volume of liquids.  $1 \text{ L} = 1000 \text{ mL}$ ,  $1000 \text{ cm}^3 = 1 \text{ dm}^3$

**Density:** Density of a substance is its amount of mass per unit volume. SI unit of density = SI unit of mass/SI unit of volume =  $\text{kg}/\text{m}^3$  or  $\text{kg m}^{-3}$ . This unit is quite large and a chemist often expresses density in  $\text{g cm}^{-3}$ .

**Temperature--** There are three common scales to measure temperature —  $^{\circ}\text{C}$  (degree celsius),  $^{\circ}\text{F}$  (degree Fahrenheit) and K (kelvin). Here, K is the SI unit.

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

$$\text{K} = ^{\circ}\text{C} + 273.15$$

Note—Temperature below  $0^{\circ}\text{C}$  (i.e. negative values) are possible in Celsius scale but in Kelvin scale, negative temperature is not possible.

### Scientific 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).

e.g. We can write 232.508 as  $2.32508 \times 10^2$  in scientific notation. Similarly, 0.00016 can be written as  $1.6 \times 10^{-4}$ .

**Precision** refers to the closeness of various measurements for the same quantity.

**Accuracy** is the agreement of a particular value to the true value of the result

### Significant Figures

The reliability of a measurement is indicated by the number of digits used to represent it. To express it more accurately we express it with digits that are known with certainty. These are called as Significant figures. They contain all the certain digits plus one doubtful digit in a number.

### Rules for Determining the Number of Significant Figures

- All non-zero digits are significant. For example, 6.9 has two significant figures, while 2.16 has three significant figures. The decimal place does not determine the number of significant figures.
- A zero becomes significant in case it comes in between non zero numbers. For example, 2.003 has four significant figures, 4.02 has three significant figures.
- Zeros at the beginning of a number are not significant. For example, 0.002 has one significant figure while 0.0045 has two significant figures.
- All zeros placed to the right of a number are significant. For example, 16.0 has three significant figures, while 16.00 has four significant figures. Zeros at the end of a number without decimal point are ambiguous.
- In exponential notations, the numerical portion represents the number of significant figures. For example, 0.00045 is expressed as  $4.5 \times 10^{-4}$  in terms of scientific notations. The number of significant figures in this number is 2, while in Avogadro's number ( $6.023 \times 10^{23}$ ) it is four.
- The decimal point does not count towards the number of significant figures. For example, the number 345601 has six significant figures but can be written in different ways, as 345.601 or 0.345601 or 3.45601 all having same number of significant figures.

### Retention of Significant Figures - Rounding off Figures

The rounding off procedure is applied to retain the required number of significant figures.

1. If the digit coming after the desired number of significant figures happens to be more than 5, the preceding significant figure is increased by one, 4.317 is rounded off to 4.32.
2. If the digit involved is less than 5, it is neglected and the preceding significant figure remains unchanged, 4.312 is rounded off to 4.31.
3. If the digit happens to be 5, the last mentioned or preceding significant figure is increased by one only in case it happens to be odd. In case of even figure, the

preceding digit remains unchanged. 8.375 is rounded off to 8.38 while 8.365 is rounded off to 8.36.

**Dimensional Analysis** During calculations generally there is a need to convert units from one system to other. This is called **factor label method** or **unit factor method** or **dimensional analysis**.

For example- 5 feet and 2 inches (height of an Indian female) is to be converted in SI unit

$$1 \text{ inch} = 2.54 \times 10^{-2} \text{ m}$$

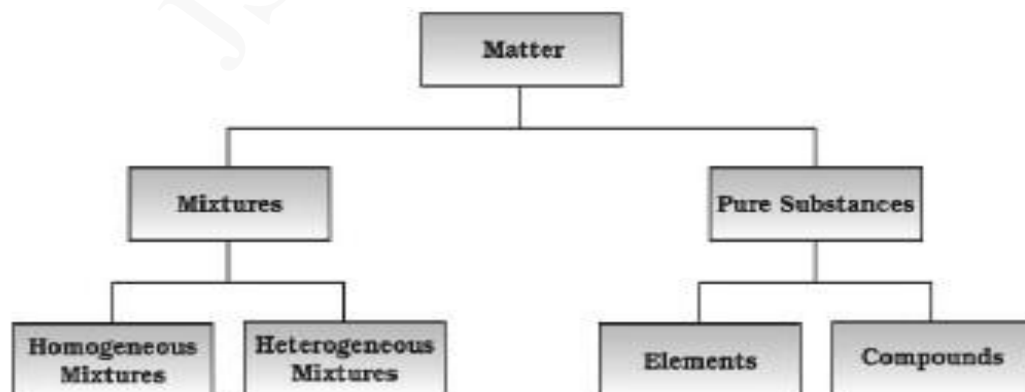
$$1 = \frac{2.54 \times 10^{-2} \text{ m}}{1 \text{ inch}} \quad \text{then, } 5 \text{ feet and } 2 \text{ inch} = 62 \text{ inch}$$

$$= 62 \text{ inch} \times \frac{2.54 \times 10^{-2} \text{ m}}{1 \text{ inch}} = 1.58 \text{ m}$$

### Physical Classification of Matter

| Properties                             | Solid  | Liquid   | Gas  |
|--|--|--|--|
| 1. volume                              | Definite   | Definite                                       | Indefinite                                     |
| 2. Shape                               | Definite   | Indefinite                                     | Indefinite                                     |
| 3. Inter molecular force of attraction | Very high  | Moderate                                       | Negligible / Very low                          |
| 4. arrangement of molecules            | Orderly arranged                                       | Free to move within the volume                 | Free to move everywhere                        |
| 5. Inter molecular space               | Very small   | Slightly greater                               | Very great                                     |
| 7. Compressibility                     | Not compressible                                       | Not compressible                               | Highly compressible                            |
| 8. Expansion on heating                | Very little  | Very little                                    | Highly expand                                  |
| 9. Rigidity                            | Very rigid   | Not rigid known as fluid                       | Not rigid and known as fluid                   |
| 9. Fluidity                            | Can't flow   | Can flow                                       | Can flow                                       |
| 10. Diffusion                          | They can diffuse due to kinetic energy of liquid/gases | Can diffuse And rate of diffusion is very fast | Can diffuse And rate of diffusion is very fast |

### Chemical Classification of matter---



## Elements

An element is the simplest form of matter that cannot be split into simpler substances or built from simpler substances by any ordinary chemical or physical method. There are 114 elements known to us, out of which 92 are naturally occurring while the rest have been prepared artificially.

Elements are further classified into metals, non-metals and metalloids.

## Compounds

A compound is a pure substance made up of two or more elements combined in a definite proportion by mass, which could be split by suitable chemical methods.

### Characteristics of compound

- Compounds always contain a definite proportion of the same elements by mass.
- The properties of compounds are totally different from the elements from which they are formed.
- Compounds are homogeneous.
- Compounds are broadly classified into inorganic and organic compounds. Inorganic compounds are those, which are obtained from non-living sources such as minerals. For example, common salt, marble and limestone. Organic compounds are those, which occur in living sources such as plants and animals. They all contain carbon. Common organic compounds are oils, wax, fats etc.

## Mixtures

A mixture is a combination of two or more elements or compounds in any proportion so that the components do not lose their identity. Air is an example of a mixture

Mixtures are of two types, homogeneous and heterogeneous.

**Homogeneous mixtures** have the same composition throughout the sample. The components of such mixtures cannot be seen under a powerful microscope. They are also called solutions. Examples of homogeneous mixtures are air, seawater, gasoline, brass etc.

**Heterogeneous mixtures** consist of two or more parts (phases), which have different compositions. These mixtures have visible boundaries of separation between the different constituents and can be seen with the naked eye e.g., sand and salt, chalk powder in water etc.

## LAWS OF CHEMICAL COMBINATIONS

**Law of Conservation of Mass** (Given by Antoine Lavoisier in 1789).

It states that matter (mass) can neither be created nor destroyed.

**Law of Definite Proportions or Law of Constant Composition:**

This law was proposed by Louis Proust in 1799, which states that:

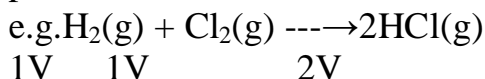
'A chemical compound always consists of the same elements combined together in the same ratio, irrespective of the method of preparation or the source from where it is taken'.

**Law of Multiple Proportions** Proposed by Dalton in 1803, this law states that:

'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 of Gaseous Volumes** (Given by Gay Lussac in 1808.)

According to this law when gases combine or are produced in a chemical reaction they do so in a simple ratio by volume provided all gases are at same temperature and pressure.



All reactants and products have simple ratio 1:1:2.

**Avogadro Law** (In 1811, Given by Avogadro)

According to this law equal volumes of gases at the same temperature and pressure should contain equal number of molecules.

**Dalton's Atomic Theory**

- All substances are made up of tiny, indivisible particles called atoms.
- Atoms of the same element are identical in shape, size, mass and other properties.
- Atoms of different elements are different in all respects.
- Atom is the smallest unit that takes part in chemical combinations.
- Atoms combine with each other in simple whole number ratios to form compound atoms called molecules.
- Atoms cannot be created, divided or destroyed during any chemical or physical change.

**Atoms and Molecules**

The smallest particle of an element, which may or may not have independent existence is called an atom, while the smallest particle of a substance which is capable of independent existence is called a molecule.

Molecules are classified as homoatomic and heteroatomic. Homoatomic molecules are made up of the atoms of the same element and heteroatomic molecules are made up of the atoms of the different element have different atomicity (number of atoms in a molecule of an element) like monoatomic, diatomic, triatomic and polyatomic.

**Atomic Mass Unit**

One atomic mass unit is defined as a mass exactly equal to one twelfth the mass of one carbon -12 atom. And  $1 \text{ amu} = 1.66056 \times 10^{-24} \text{ g}$ .

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

**Atomic Mass**

Atomic mass of an element 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.

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

**Gram Atomic Mass**

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

For example, the atomic mass of oxygen = 16 amu

Therefore gram atomic mass of oxygen = 16 g

### **Molecular Mass**

Molecular mass of a substance is defined as the average relative mass of its molecule as compared to the mass of an atom of C-12 taken as 12. It expresses as to how many times the molecule of a substance is heavier than 1/12th of the mass of an atom of carbon.

For example, a molecule of carbon dioxide is 44 times heavier than 1/12th of the mass of an atom of carbon. Therefore the molecular mass of CO<sub>2</sub> is 44 amu.

It is obtained by adding the atomic masses of all the atoms present in one molecule.

### **Gram Molecular Mass**

A quantity of substance whose mass in grams is numerically equal to its molecular mass is called gram molecular mass. In simple terms, molecular mass of a substance expressed in grams is called gram molecular mass.

e.g., the molecular mass of oxygen = 32 amu

Therefore, gram molecular mass of oxygen = 32 g

### **Formula Mass-**

Sum of atomic masses of the elements present in one formula unit of a compound. It is used for the ionic compounds.

### **Mole Concept.**

Mole is defined as the amount of a substance, which contains the same number of chemical units (atoms, molecules, ions or electrons) as there are atoms in exactly 12 grams of pure carbon-12.

A mole represents a collection of  $6.022 \times 10^{23}$  (Avogadro's number) chemical units..

**The mass of one mole of a substance in grams is called its molar mass.**

### **Molar Volume**

The volume occupied by one mole of any substance is called its molar volume. It is denoted by V<sub>m</sub>. One mole of all gaseous substances at 273 K and 1 atm pressure occupies a volume equal to 22.4 litre or 22,400 mL. The unit of molar volume is litre per mol or millilitre per mol

### **PERCENTAGE COMPOSITION—**

The mass percentage of each constituent element present in any compound is called its percentage composition

Mass % of the element =  $\frac{\text{Mass of element in 1 molecule of the compound}}{\text{Molecular mass of the compound}} \times 100$

### **Empirical Formula and Molecular Formula—**

An **empirical formula** represents the simplest whole number ratio of various atoms present in a compound. E.g. CH is the empirical formula of benzene.

The **molecular formula** shows the exact number of different types of atoms present in a molecule of a compound. E.g. C<sub>6</sub>H<sub>6</sub> is the molecular formula of benzene.

### **Relationship between empirical and molecular formulae**

The two formulas are related as Molecular formula = n x empirical formula

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

## Chemical Equation-

Shorthand representation of a chemical change in terms of symbols and formulae of the substances involved in the reaction is called chemical equation..

The substances that react among themselves to bring about the chemical changes are known as reactants, whereas the substances that are produced as a result of the chemical change, are known as products.

**Limiting Reagent-** The reactant which gets consumed first or limits the amount of product formed is known as **limiting reagent**

**Reactions in Solutions--** The concentration of a solution can be expressed in any of the following ways.

1. **Mass Percent** is the mass of the solute in grams per 100 grams of the solution.

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

A 5 % solution of sodium chloride means that 5 g of NaCl is present in 100g of the solution.

2. **Volume percent** is the number of units of volume of the solute per 100 units of the volume of solution.

$$\text{Volume \% of the solute} = \frac{\text{Volume of the solute}}{\text{Volume of the solution}} \times 100$$

A 5 % (v/v) solution of ethyl alcohol contains 5 cm<sup>3</sup> of alcohol in 100 cm<sup>3</sup> of the solution

3. **Molarity** of the solution is defined as the number of moles of solute dissolved per litre (dm<sup>3</sup>) of the solution. It is denoted by the symbol M. Measurements in Molarity can change with the change in temperature because solutions expand or contract accordingly.

$$\text{Molarity of the solution} = \frac{\text{No. of moles of the solute}}{\text{Volume of the solution in litre}} = \frac{n}{V}$$

The Molarity of the solution can also be expressed in terms of mass and molar mass

$$\text{Molarity of the solution} = \frac{\text{Mass of the solute}}{\text{Molar mass of the solute} \times \text{volume of the solution in liter}}$$

In terms of weight, molarity of the substance can be expressed as:

$$\text{Molarity} = \frac{W_g}{M \text{ g mol}^{-1} \times V \text{ litre}} = \frac{W}{M \times V} \text{ mol/L}$$

## Molarity equation

To calculate the volume of a definite solution required to prepare solution of other molarity, the following equation is used:

$M_1V_1 = M_2V_2$ , where  $M_1$ = initial molarity,  $M_2$ = molarity of the new solution,  $V_1$ = initial volume and  $V_2$ = volume of the new solution.

4. **Molality-** Molality is defined as the number of moles of solute dissolved per 1000 g (1 kg) of solvent. Molality is expressed as 'm'.

$$\text{Molality} = \frac{\text{Moles of the solute}}{\text{Wt. of Solvent (in gm)}} \times 1000$$



5. Mole Fraction is the ratio of number of moles of one component to the total number of moles (solute and solvents) present in the solution. It is expressed as 'x'.

$$\text{Mole fraction of the solute} = \frac{\text{Moles of the solute}}{\text{Moles of solute} + \text{Moles of solvent}}$$

$$\text{Mole fraction of the solvent} = \frac{\text{Moles of the solvent}}{\text{Moles of solute} + \text{Moles of solvent}}$$

$$\text{Mole fraction of the solute} + \text{Mole fraction of solvent} = 1$$

**One Mark questions with answers**

1. What is the significant figures in  $1.050 \times 10^4$ ?

Ans. Four

2. What is the S.I. unit of Density?

Ans.  $\text{Kg m}^{-3}$

3. What do mean by Mole fraction?

Ans. Mole Fraction is the ratio of number of moles of one component to the total number of moles (solute and solvents) present in the solution. It is expressed as 'x'.

4. Round off up to 3 significant figure (a) 1.235      (b) 1.225

Ans. (a) 1.24    (b) 1.22

5. What is AZT?

Ans. Azidothymidine.

6. What is limiting reagent?

Ans. The reactant which gets consumed first or limits the amount of product formed is known as **limiting reagent**

7. What is the relation between temperature in degree Celsius and degree fahrenheit?

Ans.

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

8. Define one mole?

Ans. One mole is the amount of a substance that contains as many particles as there are atoms in exactly 12 g of the carbon-12.

9. Calculate the formula mass calcium chloride.

Ans. Formula mass of  $\text{CaCl}_2 = 40 + 2 \times 35.5 = 40 + 71 = 111 \text{ u}$

10. What is the law called which deals with the ratios of the volumes of the gaseous reactants and products?

Ans. Gay Lussac's law of gaseous volumes.

**Two Marks questions with answers**

1. Give the two points of differences between homogeneous and heterogeneous mixtures.

Ans.

| Homogeneous mixture  | Heterogeneous mixture  |
|--|--|
| 1. Homogeneous mixtures have the same composition throughout the sample.<br><br>2. The components of such mixtures cannot be seen under a powerful microscope. | 1. Heterogeneous mixtures consist of two or more parts (phases), which have different compositions.<br><br>2. These mixtures have visible boundaries of separation between the different constituents and can be seen with the naked eye |

2. Copper oxide obtained by heating copper carbonate or copper nitrate contains copper and oxygen in the same ratio by mass. Which law is illustrated by this observation? State the law.

**Ans. Law of Definite Proportions** This law states that: A chemical compound always consists of the same elements combined together in the same ratio, irrespective of the method of preparation or the source from where it is taken.

3. Write the empirical formula of the following:

(a)  $N_2O_4$       (b)  $C_6H_{12}O_6$       (c)  $H_2O$       (d)  $H_2O_2$

Ans. (a)  $NO_2$       (b)  $CH_2O$       (c)  $H_2O$       (d)  $HO$

4. Briefly explain the difference between precision and accuracy.

Ans. 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.

5. Define the law of multiple proportions. Explain it with one example.

Ans. 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. For example- carbon combines with oxygen to form two compounds  $CO$  and  $CO_2$ .

| Compound  | $CO$ | $CO_2$ |
|-----------|------|--------|
| Mass of C | 12   | 12     |
| Mass of O | 16   | 32     |

Masses of oxygen which combine with a fixed mass of carbon (12g) bear a simple ratio of 16:32 or 1:2.

6. Chlorine has two isotopes of atomic mass units 34.97 and 36.97. The relative abundance of the isotopes is 0.755 and 0.245 respectively. Find the average atomic mass of chlorine.

Ans. Average atomic mass =  $34.97 \times 0.755 + 36.97 \times 0.245 = 35.46 \text{ u}$

7. Calculate the percentage composition of water.

Ans. Mass % of an element =  $\frac{\text{mass of that element in the compound} \times 100}{\text{molar mass of the compound}}$

Molar mass of water = 18.02 g

Mass % of hydrogen =  $\frac{2 \times 1.008 \times 100}{18.02}$

= 11.18

Mass % of oxygen =  $\frac{16.00 \times 100}{18.02} = 88.79$

8. State the number of significant figures in each of the following:

- (i) 208.91 (ii) 0.00456 (iii) 453 (iv) 0.346

Ans.

(i) 208.91 has five significant figures.

(ii) 0.00456 has three significant figures.

(iii) 453 has three significant figures.

(iv) 0.346 has three significant figures.

8. Express the results of the following calculations to the appropriate number of significant figures.

(i)  $\frac{3.24 \times 0.08666}{5.006}$  (ii)  $\frac{(1.36 \times 10^{-4})(0.5)}{2.6}$

Ans.

(i)  $\frac{3.24 \times 0.08666}{5.006} = 0.05608 = 0.0561$

(ii)  $\frac{(1.36 \times 10^{-4})(0.5)}{2.6} = 0.2615 \times 10^{-4} = 0.3 \times 10^{-4}$

9. How are 0.50 mol Na<sub>2</sub>CO<sub>3</sub> and 0.50 M Na<sub>2</sub>CO<sub>3</sub> different?

Ans. Molar mass of Na<sub>2</sub>CO<sub>3</sub> = 2 x 23 + 12 + 3 x 16 = 106 g / mol

0.50 mol Na<sub>2</sub>CO<sub>3</sub> means 0.50 x 106 = 53 g

0.50 M Na<sub>2</sub>CO<sub>3</sub> means 0.50 mol i.e. 53 g of Na<sub>2</sub>CO<sub>3</sub> are present in 1 L of the solution.

### Three Marks questions with answers-

1. What is unit factor method? Express the following in SI units - 93 million miles (distance between earth and sun)

Ans. Method to convert units from one system to other is called unit factor method.

93 million miles = 93 x 10<sup>6</sup> miles

1 mile = 1.60934 km = 1.60934 x 10<sup>3</sup> m

$$1 = \frac{1.60934 \times 10^3 \text{ m}}{1 \text{ mile}}$$

$$93 \text{ million mile} = 93 \times 10^6 \text{ mile} \times \frac{1.60934 \times 10^3 \text{ m}}{1 \text{ mile}}$$

$$= 1.5 \times 10^{11} \text{ m}$$

2. Write the three points of difference between compound and mixture.

Ans.

| Compound   | Mixture                                  |
|--|--|
| Constituents are always present in a fixed ratio by mass | Constituents may be present in any ratio |
| May or may not be homogeneous in nature                  | Always homogeneous in nature             |
| Constituents can be easily separated                     | Constituents cannot be easily            |

|  |   |
|--|---|
| by simple mechanical means                               | separated by simple mechanical means                              |
| Properties are midway between those of its constituents. | Properties are entirely different from those of its constituents. |

3. What do mean by gram atomic mass. One million silver atoms weigh  $1.79 \times 10^{16}$  g. Calculate the gram atomic mass of silver.

Ans. atomic mass of an element expressed in grams is the gram atomic mass

Number of silver atoms = 1 million =  $1 \times 10^6$   
 Mass of one million Ag atoms =  $1.79 \times 10^{16}$  g

$$\begin{aligned} \text{Mass of } 6.023 \times 10^{23} \text{ atoms of silver} &= \frac{1.79 \times 10^{16} \text{ g} \times 6.023 \times 10^{23}}{1 \times 10^6} \\ &= 107.8 \text{ g} \end{aligned}$$

4. What is the percentage of carbon, hydrogen and oxygen in ethanol?

Ans. Molecular formula of ethanol is :  $\text{C}_2\text{H}_5\text{OH}$

Molar mass of ethanol is :  $(2 \times 12.01 + 6 \times 1.008 + 16.00) \text{ g} = 46.068 \text{ g}$

Mass per cent of carbon =  $(24.02 \text{ g} / 46.068 \text{ g}) \times 100 = 52.14\%$

Mass per cent of hydrogen =  $(6.048 \text{ g} / 46.068 \text{ g}) \times 100 = 13.13\%$  Mass per cent of oxygen =  $(16.00 \text{ g} / 46.068 \text{ g}) \times 100 = 34.73\%$

5. What do mean by molarity .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. The number of moles of solute dissolved per litre ( $\text{dm}^3$ ) of the solution is called molarity

Since molarity (M) = No. of moles of solute / Volume of solution in litres

$$= (\text{Mass of NaOH} / \text{Molar mass of NaOH}) / 0.250 \text{ L}$$

$$= (4 \text{ g} / 40 \text{ g}) / 0.250 \text{ L} = 0.1 \text{ mol} / 0.250 \text{ L}$$

$$= 0.4 \text{ mol L}^{-1}$$

$$= 0.4 \text{ M}$$

6. Classify the following as pure substances or mixture-

(a) ethyl alcohol (b) oxygen (c) blood (d) carbon (e) steel (f) distilled water

Ans. Pure substance- ethyl alcohol , oxygen , carbon, distilled water

Mixture- blood, steel

7. What are the rules for rounding off?

Ans. 1. If the digit coming after the desired number of significant figures happens to be more than 5, the preceding significant figure is increased by one,

2. If the digit involved is less than 5, it is neglected and the preceding significant figure remains unchanged,

3. If the digit happens to be 5, the last mentioned or preceding significant figure is increased by one only in case it happens to be odd. In case of even figure, the preceding digit remains unchanged.
8. Define –(a) Average atomic mass (b) Molecular mass (c) Formula mass  
 Ans. (a) Average atomic mass- Atomic mass of an element 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.  
 (b) Molecular mass- it is sum of atomic masses of the elements present in a molecule.  
 (c) Formula mass- it is sum of atomic masses of the elements present in a formula unit of a compound.
9. Express the following in the scientific notation with 2 significant figures-  
 (a) 0.0048 (b) 234,000 (c) 200.0  
 Ans. (a)  $4.8 \times 10^{-3}$  (b)  $2.3 \times 10^5$  (c)  $2.0 \times 10^2$
10. Calculate the number of atoms in each of the following (i) 52 moles of Ar  
 (ii) 52 u of He (iii) 52 g of He  
 Ans. (i) 1 mole of Ar =  $6.022 \times 10^{23}$  atoms of Ar  
 $\therefore$  52 mol of Ar =  $52 \times 6.022 \times 10^{23}$  atoms of Ar  
 =  $3.131 \times 10^{25}$  atoms of Ar  
 (ii) 1 atom of He = 4 u of He  
 Or,  
 4 u of He = 1 atom of He  
  
 1 u of He =  $1/4$  atom of He  
 52u of He =  $52/4$  atom of He  
 = 13 atoms of He  
 (iii) Molar mass of He = 4 g/mol  
 4 g of He contains =  $6.022 \times 10^{23}$  atoms of He  
 52 g of He contains =  $\frac{6.022 \times 10^{23} \times 52}{4}$  =  $78.286 \times 10^{23}$  atoms of He

**Five Marks questions with answers-**

1. What is the difference between empirical and molecular formula? 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. An empirical formula represents the simplest whole number ratio of various atoms present in a compound whereas the molecular formula shows the exact number of different types of atoms present in a molecule of a compound.

| Name of element | Percentage of elements | Step-1 Conversion of mass per cent to grams. | Step 2. number moles of each element | Step 3. Divide the mole value by the smallest number |
|-----------------|------------------------|--|--------------------------------------|--|
| C               | 24.27%                 | 24.27g                                       | $24.27/12 = 2.0225$                  | $2.0225/2.018 = 1$                                   |
| H               | 4.07%                  | 4.07g  | $4.07/1 = 4.07$                      | $4.07/2.018 = 2$                                     |
| Cl              | 71.65%                 | 71.65g                                       | $71.65/35.5 = 2.018$                 | $2.018/2.018 = 1$                                    |

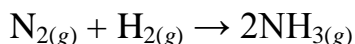
The empirical formula of the above compound is CH<sub>2</sub>Cl.

empirical formula mass is  $12 + (1 \times 2) + 35.5 = 49.5$

$n = \text{molecular mass} / \text{empirical formula mass} = 98.96 / 49.5 = 2$

Hence molecular formula is C<sub>2</sub>H<sub>4</sub>Cl<sub>2</sub>

2. Dinitrogen and dihydrogen react with each other to produce ammonia according to the following chemical equation:



(i) Calculate the mass of ammonia produced if  $2.00 \times 10^3$  g dinitrogen reacts with  $1.00 \times 10^3$  g of dihydrogen.

(ii) Will any of the two reactants remain unreacted?

(iii) If yes, which one and what would be its mass?

Ans. (i) Balancing the given chemical equation,  $\text{N}_{2(g)} + 3\text{H}_{2(g)} \longrightarrow 2\text{NH}_{3(g)}$

From the equation, 1 mole (28 g) of dinitrogen reacts with 3 mole (6 g) of dihydrogen to give 2 mole (34 g) of ammonia.

$\Rightarrow 2.00 \times 10^3$  g of dinitrogen will react with  $\frac{6 \text{ g}}{28 \text{ g}} \times 2.00 \times 10^3 \text{ g}$  dihydrogen i.e.,

$2.00 \times 10^3$  g of dinitrogen will react with 428.6 g of dihydrogen.

Given,

Amount of dihydrogen =  $1.00 \times 10^3$  g

Hence, N<sub>2</sub> is the limiting reagent.

$\therefore$  28 g of N<sub>2</sub> produces 34 g of NH<sub>3</sub>.

Hence, mass of ammonia produced by 2000 g of N<sub>2</sub>  $= \frac{34 \text{ g}}{28 \text{ g}} \times 2000 \text{ g}$

= 2428.57 g

(ii)  $N_2$  is the limiting reagent and  $H_2$  is the excess reagent. Hence,  $H_2$  will remain unreacted.

$$\begin{aligned} \text{(iii) Mass of dihydrogen left unreacted} &= 1.00 \times 10^3 \text{ g} - 428.6 \text{ g} \\ &= 571.4 \text{ g} \end{aligned}$$

3. A welding fuel gas contains carbon and hydrogen only. Burning a small sample of it in oxygen gives 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. Calculate (i) empirical formula, (ii) molar mass of the gas, and (iii) molecular formula.

Ans. (i) 1 mole (44 g) of  $CO_2$  contains 12 g of carbon.

$$\begin{aligned} 3.38 \text{ g of } CO_2 \text{ will contain carbon} &= \frac{12 \text{ g}}{44 \text{ g}} \times 3.38 \text{ g} \\ &= 0.9217 \text{ g} \end{aligned}$$

18 g of water contains 2 g of hydrogen.

$$\begin{aligned} 0.690 \text{ g of water will contain hydrogen} &= \frac{2 \text{ g}}{18 \text{ g}} \times 0.690 \\ &= 0.0767 \text{ g} \end{aligned}$$

Since carbon and hydrogen are the only constituents of the compound, the total mass of the compound is:  $= 0.9217 \text{ g} + 0.0767 \text{ g} = 0.9984 \text{ g}$

$$\begin{aligned} \text{Percent of C in the compound} &= \frac{0.9217 \text{ g}}{0.9984 \text{ g}} \times 100 \\ &= 92.32\% \end{aligned}$$

$$\begin{aligned} \text{Percent of H in the compound} &= \frac{0.0767 \text{ g}}{0.9984 \text{ g}} \times 100 \\ &= 7.68\% \end{aligned}$$

$$\begin{aligned} \text{Moles of carbon in the compound} &= \frac{92.32}{12.00} \\ &= 7.69 \end{aligned}$$

$$\begin{aligned} \text{Moles of hydrogen in the compound} &= \frac{7.68}{1} \\ &= 7.68 \end{aligned}$$

Ratio of carbon to hydrogen in the compound  $= 7.69 : 7.68 = 1 : 1$

Hence, the empirical formula of the gas is CH.

(ii) Given,

Weight of 10.0L of the gas (at S.T.P) = 11.6 g

$$\begin{aligned} \text{Weight of 22.4 L of gas at STP} &= \frac{11.6 \text{ g}}{10.0 \text{ L}} \times 22.4 \text{ L} \\ &= 25.984 \text{ g} \approx 26 \text{ g} \end{aligned}$$

Hence, the molar mass of the gas is 26 g.

(iii) Empirical formula mass of CH = 12 + 1 = 13 g

$$n = \frac{\text{Molar mass of gas}}{\text{Empirical formula mass of gas}}$$

$$= \frac{26 \text{ g}}{13 \text{ g}}$$

$$n = 2$$

∴ Molecular formula of gas = (CH)<sub>n</sub> = C<sub>2</sub>H<sub>2</sub>

### HOTS (Higher Order Thinking Skills)

1. What is the difference between 160 cm and 160.0 cm

Ans. 160 has three significant figures while 160.0 has four significant figures. Hence, 160.0 represents greater accuracy.

2. In the combustion of methane, what is the limiting reactant and why?

Ans. Methane is the limiting reactant because the other reactant is oxygen of the air which is always present in excess. Thus, the amounts of CO<sub>2</sub> and H<sub>2</sub>O formed depend upon the amount of methane burnt.

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

a. Atomic masses of the elements A and B

b. Molecular formula of the compound, if its molecular mass is found to be 160

Ans. Relative no. of moles of an element =  $\frac{\% \text{ of the element}}{\text{Atomic mass}}$

$$\text{Or atomic mass} = \frac{\% \text{ of the element}}{\text{Relative no. of moles}} = \frac{70}{1.25} = 56$$

$$\text{Atomic mass of B} = \frac{30}{1.88} = 16$$

Calculation of Empirical formula

| Element | Relative no. of moles | Simplest molar ratio | Simplest whole no. molar ratio |
|---------|-----------------------|----------------------|--------------------------------|
| A       | 1.25                  | 1.25/1.25 = 1        | 2                              |
| B       | 1.88                  | 1.88/1.25 = 1.5      | 3                              |

Empirical formula = A<sub>2</sub>B<sub>3</sub>

Calculation of molecular formula-

$$\text{Empirical formula mass} = 2 \times 56 + 3 \times 16 = 160$$

$$n = \frac{\text{molecular mass}}{\text{Empirical formula mass}} = \frac{160}{160} = 1$$

Molecular formula = A<sub>2</sub>B<sub>3</sub>