1 Some Basic Concepts of Chemistry

Facts that Matter

• Importance of Chemistry

Chemistry has a direct impact on our life and has wide range of applications in different fields. These are given below:

(A) In Agriculture and Food:

(*i*) It has provided chemical fertilizers such as urea, calcium phosphate, sodium nitrate, ammonium phosphate etc.

(*ii*) It has helped to protect the crops from insects and harmful bacteria, by the use of certain effective insecticides, fungicides and pesticides.

(*iii*) The use of preservatives has helped to preserve food products like jam, butter, squashes etc. for longer periods.

(B) In Health and Sanitation:

(*i*) It has provided mankind with a large number of life-saving drugs. Today, dysentery and pneumonia are curable due to discovery of sulpha drugs and penicillin life-saving drugs. Cisplatin and taxol have been found to be very effective for cancer therapy and AZT (Azidothymidine) is used for AIDS victims.

(*ii*) Disinfectants such as phenol are used to kill the micro-organisms present in drains, toilet, floors etc.

(*iii*) A low concentration of chlorine *i.e.,* 0.2 to 0.4 parts per million (ppm) is used for sterilization of water to make it fit for drinking purposes.

(C) Saving the Environment:

The rapid industrialisation all over the world has resulted in lot of pollution. Poisonous gases and chemicals are being constantly released in the atmosphere. They are polluting environment at an alarming rate. Scientists are working day and night to develop substitutes which may cause lower pollution. For example, CNG (Compressed Natural Gas), a substitute of petrol, is very effective in checking pollution caused by automobiles.

(D) Application in Industry:

Chemistry has played an important role in developing many industrially manufactured fertilizers, alkalis, acids, salts, dyes, polymers, drugs, soaps, detergents, metal alloys and other inorganic and organic chemicals including new materials contribute in a big way to the national economy.

• Matter

Anything which has mass and occupies space is called matter.

For example, book, pencil, water, air are composed of matter as we know that they have mass and they occupy space.

• Classification of Matter

There are two ways of classifying the matter:

(A) Physical classification

(B) Chemical classification

(A) Physical Classification:

Matter can exist in three physical states:

1. Solids **2.** Liquids **3.** Gases

1. **Solids:** The particles are held very close to each other in an orderly fashion and there is not much freedom of movement.
   Characteristics of solids: Solids have definite volume and definite shape.
2. **Liquids:** In liquids, the particles are close to each other but can move around.
   Characteristics of liquids: Liquids have definite volume but not definite shape.
3. **Gases:** In gases, the particles are far apart as compared to those present in solid or liquid states. Their movement is easy and fast.
   Characteristics of Gases: Gases have neither definite volume nor definite shape. They completely occupy the container in which they are placed.

(B) Chemical Classification:

Based upon the composition, matter can be divided into two main types:

1. Pure Substances **2.** Mixtures.

1. Pure substances: A pure substance may be defined as a single substance (or matter) which cannot be separated by simple physical methods.

Pure substances can be further classified as

(*i*) Elements (*ii*) Compounds

(*i*) **Elements:** An element consists of only one type of particles. These particles may be atoms or molecules.

For example, sodium, copper, silver, hydrogen, oxygen etc. are some examples of elements. They all contain atoms of one type. However, atoms of different elements are different in nature. Some elements such as sodium or copper contain single atoms held together as their constituent particles whereas in some others two or more atoms combine to give molecules of the element. Thus, hydrogen, nitrogen and oxygen gases consist of molecules in which two atoms combine to give the respective molecules of the element.

(*ii*) **Compounds:** It may be defined as a pure substance containing two or more elements combined together in a fixed proportion by weight and can be decomposed into these elements by suitable chemical methods.

Moreover, the properties of a compound are altogether different from the constituting elements.

The compounds have been classified into two types. These are:

(*i*) **Inorganic Compounds:** These are compounds which are obtained from non-living sources such as rocks and minerals. A few examples are:
Common salt, marble, gypsum, washing soda etc.

(*ii*) **Organic Compounds** are the compounds which are present in plants and animals. All the organic compounds have been found to contain carbon as their essential constituent. *For example,* carbohydrates, proteins, oils, fats etc.

2. Mixtures: The combination of two or more elements or compounds which are not chemically combined together and may also be present in any proportion, is called mixture. A few examples of mixtures are: milk, sea water, petrol, lime water, paint glass, cement, wood etc.

Types of mixtures: Mixtures are of two types:

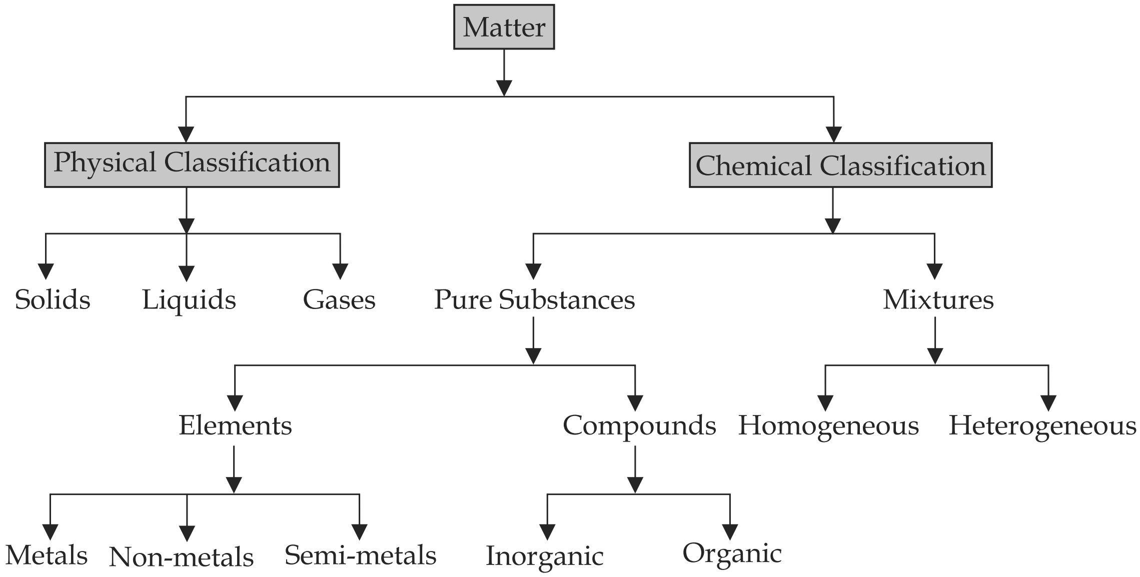
(*i*) **Homogeneous mixtures:** A mixture is said to be homogeneous if it has a uniform composition throughout and there are no visible boundaries of separation between the constituents.
For example: A mixture of sugar solution in water has the same sugar water composition throughout and all portions have the same sweetness.

(*ii*) **Heterogeneous mixtures:** A mixture is said to be heterogeneous if it does not have uniform composition throughout and has visible boundaries of separation between the various constituents. The different constituents of a heterogeneous mixture can be seen even with naked eye.
For example: When iron filings and sulphur powder are mixed together, the mixture formed is heterogeneous. It has greyish-yellow appearance and the two constituents, iron and sulphur, can be easily identified with naked eye.

• Differences between Compounds and Mixtures

Compounds	*Mixtures*
1. In a compound, two or more elements are combined chemically.	1. In a mixture, two or more elements or compounds are simply mixed and not combined chemically.
2. In a compound, the elements are present in the fixed ratio by mass. This ratio cannot change.	2. In a mixture the constituents are not present in fixed ratio. It can vary.
3. Compounds are always homogeneous *i.e.,* they have the same composition throughout.	3. Mixtures may be either homogeneous or heterogeneous in nature.
4. In a compound, constituents cannot be separated by physical methods.	4. Constituents of mixtures can be separated by physical methods.
5. In a compound, the constituents lose their identities *i.e.,* a compound does not show the characteristics of the constituting elements.	5. In a mixture, the constituents do not lose their identities *i.e.,* a mixture shows the characteristics of all the constituents.

We have discussed the physical and chemical classification of matter. A flow sheet representation of the same is given below.



- **Properties of Matter and Their Measurements**

Physical Properties: Those properties which can be measured or observed without changing the identity or the composition of the substance.

Some examples of physical properties are colour, odour, melting point, boiling point etc.

Chemical Properties: It requires a chemical change to occur. The examples of chemical properties are characteristic reactions of different substances. These include acidity, basicity, combustibility etc.

- **Units of Measurement**

Fundamental Units: The quantities mass, length and time are called fundamental quantities and their units are known as fundamental units.

There are seven basic units of measurement for the quantities: length, mass, time, temperature, amount of substance, electric current and luminous intensity.

SI-System: This system of measurement is the most common system employed throughout the world.

It has given units of all the seven basic quantities listed above.

Table 1.1. Basic Physical Quantities and their Units

Basic Physical Quantity	*Symbol for Quantity*	*Name of SI Unit*	*Symbol for SI Unit*
Length	l	metre	m
Mass	m	kilogram	kg
Time	t	second	s
Electric current	I	ampere	A
Thermodynamic temperature	T	kelvin	K
Amount of substance	n	mole	mol
Luminous intensity	I_v	candela	cd

- **Definitions of Basic SI Units**

1. **Metre:** It is the length of the path travelled by light in vacuum during a time interval of 1/299792458 of a second.
2. **Kilogram:** It is the unit of mass. It is equal to the mass of the international prototype of the kilogram.
3. **Second:** It is the duration of 9192631, 770 periods of radiation which correspond to the transition between the two hyper fine levels of the ground state of caesium-133 atom.
4. **Kelvin:** It is the unit of thermodynamic temperature and is equal to 1/273.16 of the thermodynamic temperature of the triple point of water.
5. **Ampere:** The ampere is that constant current which if maintained in two straight parallel conductors of infinite length, of negligible circular cross section and placed, 1 metre apart in vacuum, would produce between these conductors a force equal to 2×10^{-7} N per metre of length.
6. **Candela:** It may be defined as the luminous intensity in a given direction, from a source which emits monochromatic radiation of frequency 540×10^{12} Hz and that has a radiant intensity in that direction of 1/683 watt per steradian.
7. **Mole:** It is the amount of substance which contains as many elementary entities as there are atoms in 0.012 kilogram of carbon –12. Its symbol is 'mol'.

- **Mass and Weight**

Mass: Mass of a substance is the amount of matter present in it.
The mass of a substance is constant.
The mass of a substance can be determined accurately in the laboratory by using an analytical balance.
SI unit of mass is kilogram.

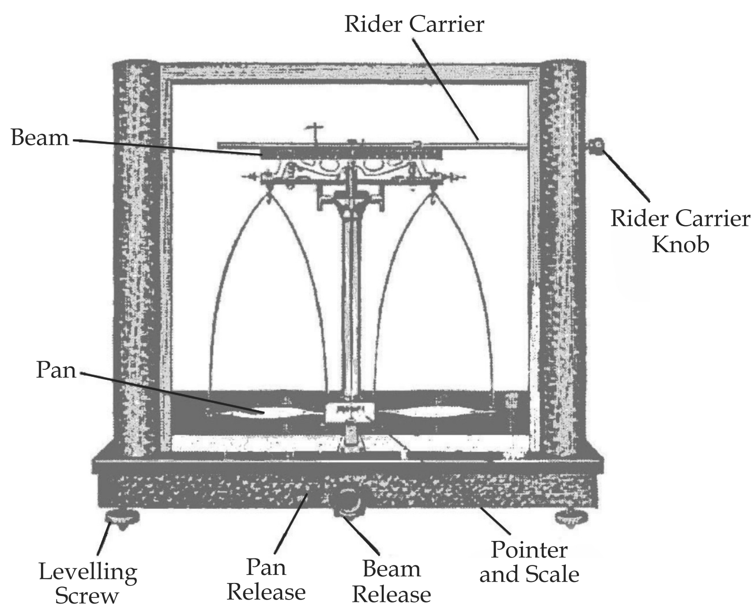


Fig. 1.1 *Analytical balance.*

Weight: It is the force exerted by gravity on an object. Weight of substance may vary from one place to another due to change in gravity.

Volume: Volume means the space occupied by matter. It has the units of $(\text{length})^3$. In SI units, volume is expressed in metre^3 (m^3). However, a popular unit of measuring volume, particularly in liquids is litre (L) but it is not in SI units or an S.I. unit.

Mathematically,

$$1\text{L} = 1000\ \text{mL} = 1000\ \text{cm}^3 = 1\text{dm}^3.$$

Volume of liquids can be measured by different devices like burette, pipette, cylinder, measuring flask etc. All of them have been calibrated.

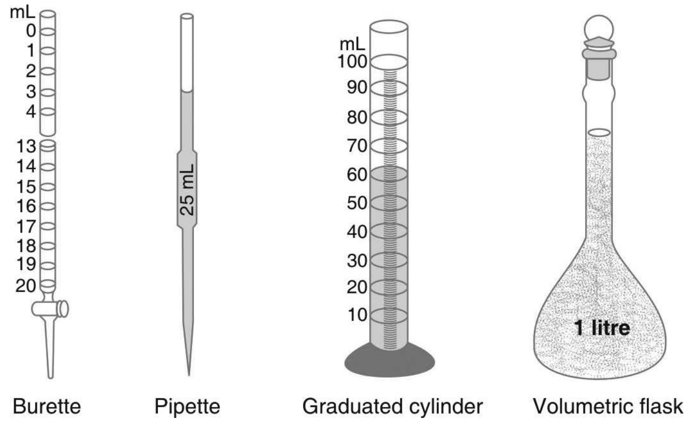


Fig. 1.2 *Some volume measuring devices.*

Temperature: There are three scales in which temperature can be measured. These are known as Celsius scale (°C), Fahrenheit scale (°F) and Kelvin scale (K).

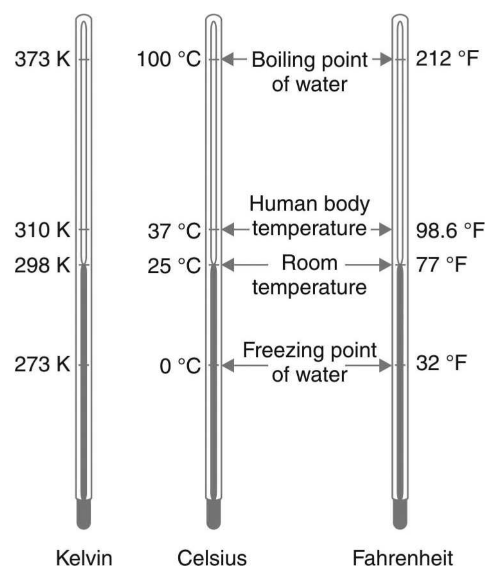


Fig. 1.3 *Thermometers using different temperature scales.*

→ Thermometres with Celsius scale are calibrated from 0°C to 100°C.
→ Thermometres with Fahrenheit scale are calibrated from 32°F to 212°F.
→ Kelvin scale of temperature is S.I. scale and is very common these days. Temperature on this scale is shown by the sign K.

The temperature on two scales are related to each other by the relationship

$$\boxed{°F = \frac{9}{5}(°C) + 32}$$

The Kelvin scale is related to Celsius scale as follows:

$$\boxed{K = °C + 273.15}$$

Density: Density of a substance is its amount of mass per unit volume. So, SI unit of density can be obtained as follows:

$$\text{SI unit of density} = \frac{\text{SI unit of mass}}{\text{SI unit of volume}} = \frac{kg}{m^3} \text{ or } kgm^{-3}$$

This unit is quite large and a chemist often expresses density in $g\,cm^{-3}$ where mass is expressed in gram and volume is expressed in cm^3.

• Uncertainty in Measurements

All scientific measurements involve certain degree of error or uncertainty. The errors which arise depend upon two factors.

(*i*) Skill and accuracy of the worker (*ii*) Limitations of measuring instruments.

• Scientific Notation

It is an exponential 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. Thus, 232.508 can be written as 2.32508×10^2 in scientific notation.

Now let us see how calculations are carried out with numbers expressed in scientific notation.

(*i*) **Calculation involving multiplication and division**

(*a*) $(5.7 \times 10^6) \times (4.2 \times 10^5) = (5.7 \times 4.2)(10^{6+5}) = 23.94 \times 10^{11}$

(*b*) $(5.7 \times 10^6) \div (4.2 \times 10^3)$

$$\frac{5.7 \times 10^6}{4.2 \times 10^3} = \frac{5.7}{4.2} \times 10^{6-3} = 1.357 \times 10^3$$

(*ii*) **Calculation involving addition and subtraction:** For these two operations, the first numbers are written in such a way that they have the same exponent. After that, the coefficients are added or subtracted as the case may be. For example,

(*a*) $4.56 \times 10^3 + 2.62 \times 10^2$

$$= 45.6 \times 10^2 + 2.62 \times 10^2$$
$$= (45.6 + 2.62) \times 10^2$$
$$= 48.22 \times 10^2$$

(*b*) $4.5 \times 10^{-3} - 2.6 \times 10^{-4}$

$$= 4.5 \times 10^{-3} - 0.26 \times 10^{-3}$$
$$= (4.5 - 0.26) \times 10^{-3}$$
$$= 4.24 \times 10^{-3}$$

• Significant Figures

Significant figures are meaningful digits which are known with certainty. There are certain rules for determining the number of significant figures. These are stated below:

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 zeros indicates the position of decimal point.
   For example, 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 provided they are on the right side of the decimal point. For example, 0.200 g has three significant figures.
5. Counting numbers of objects. For example, 2 balls or 20 eggs have infinite significant figures as these are exact numbers and can be represented by writing infinite number of zeros after placing a decimal.
   i.e., $2 = 2.000000$
   or $20 = 20.000000$

• Addition and Subtraction of Significant Figures

In addition or subtraction of the numbers having different precisions, the final result should be reported to the same number of decimal places as in the term having the least number of decimal places.

For example, let us carry out the addition of three numbers 3.52, 2.3 and 6.24, having different precisions or different number of decimal places.

$$\begin{array}{r} 3.52 \\ 2.3 \\ 6.24 \\ \hline 12.06 \\ \hline \end{array}$$

The final result has two decimal places but the answer has to be reported only upto one decimal place, *i.e.*, the answer would be 12.0.

Subtraction of numbers can be done in the same way as the addition.

$$\begin{array}{r} 23.4730 \\ 12.11 \\ \hline 11.3630 \\ \hline \end{array}$$

The final result has four decimal places. But it has to be reported only up to two decimal places, *i.e.*, the answer would be 11.36.

• Multiplication and Division of Significant Figures

In the multiplication or division, the final result should be reported upto the same number of significant figures as present in the least precise number.

Multiplication of Numbers: $2.2120 \times 0.011 = 0.024332$

According to the rule the final result = 0.024

Division of Numbers: $4.2211 \div 3.76 = 1.12263$

The correct answer = 1.12

• Dimensional Analysis

Often while calculating, there is a need to convert units from one system to other. The method used to accomplish this is called factor label method or unit factor method or dimensional analysis.

Example. *A jug contains 2L of milk. Calculate the volume of the milk in m^3.*

Solution. Since $1 \text{ L} = 1000 \text{ cm}^3$

and 1 m = 100 cm which gives

$$\frac{1 \text{ m}}{100 \text{ cm}} = 1 = \frac{100 \text{ cm}}{1 \text{ m}}$$

To get m^3 from the above unit factors, the first unit factor is taken and it is cubed.

$$\left(\frac{1 \text{ m}}{100 \text{ cm}}\right)^3 = \frac{1 \text{ m}^3}{10^6 \text{ cm}^3} = (1)^3 = 1$$

Now, $2 \text{ L} = 2 \times 1000 \text{ cm}^3$

The above is multiplied by the unit factor

$$2 \times 1000 \text{ cm}^3 \times \frac{1 \text{ m}^3}{10^6 \text{ cm}^3} = \frac{2 \text{ m}^3}{10^3} = 2 \times 10^{-3} \text{ m}^3.$$

• Laws of Chemical Combinations

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

(*i*) Law of Conservation of Mass
(*ii*) Law of Definite Proportions
(*iii*) Law of Multiple Proportions
(*iv*) Law of Gaseous Volume (Gay Lussac's Law)
(*v*) Avogadro’s Law

(*i*) **Law of Conservation of Mass**

The law was established by a French chemist, A. Lavoisier. The law states:

In all physical and chemical changes, the total mass of the reactants is equal to that of the products.

In other words, matter can neither be created nor destroyed.

The following experiments illustrate the truth of this law.

(*a*) **When matter undergoes a physical change.**

$$\text{Ice} \xrightarrow{\text{Heat}} \text{water}$$

It is found that there is no change in weight though a physical change has taken place.

(*b*) **When matter undergoes a chemical change.**

For example, decomposition of mercuric oxide.

$$\underset{100 \text{ g}}{HgO\,(s)} \longrightarrow \underset{92.6 \text{ g}}{Hg(l)} + \underset{7.4 \text{ g}}{\tfrac{1}{2}\,O_2\,(g)}$$

During the above decomposition reaction, matter is neither gained nor lost.

(*ii*) **Law of Definite Proportions**

According to this law:

A pure chemical compound always consists of the same elements combined together in a fixed proportion by weight.

For example, Carbon dioxide may be formed in a number of ways *i.e.,*

(*a*) **By burning coke in air**

$$\underset{\text{Coke}}{C(s)} + \underset{\text{Air}}{O_2(g)} \xrightarrow{\text{Heat}} CO_2(g)$$

(*b*) **By the decomposition of limestone ($CaCO_3$) on heating**

$$\underset{\text{Limestone}}{CaCO_3(s)} \xrightarrow{\text{Heat}} \underset{\text{Calcium oxide}}{CaO(s)} + CO_2(g)$$

(*c*) **By the action of dil HCl on washing soda (Na_2CO_3)**

$$\underset{\text{Sodium carbonate}}{Na_2CO_3(s)} + 2HCl(g) \longrightarrow 2NaCl(s) + H_2O(l) + CO_2(g)$$

In all the three samples of CO_2 carbon and oxygen are in the ratio 3 : 8 by weight.

(*iii*) **Law of Multiple Proportions**

If two elements combine to form two or more compounds, the weight of one of the elements which combines with a fixed weight of the other in these compounds, bears simple whole number ratio by weight.

For example,

$$\underset{2g}{\text{Hydrogen}} + \underset{16g}{\text{Oxygen}} \longrightarrow \underset{18g}{\text{Water}}$$

$$\underset{2g}{\text{Hydrogen}} + \underset{32g}{\text{Oxygen}} \longrightarrow \underset{34g}{\text{Hydrogen Peroxide}}$$

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.

(*iv*) G**ay Lussac's Law of Gaseous Volumes**

The law states that, under similar conditions of temperature and pressure, whenever gases combine, they do so in volumes which bear simple whole number ratio with each other and also with the gaseous products.

The law may be illustrated by the following examples.

(*a*) **Combination between hydrogen and chlorine:**

$$\underset{\text{1 Volume}}{H_2(g)} + \underset{\text{1 Volume}}{Cl_2(g)} \longrightarrow \underset{\text{2 Volumes}}{2HCl\ (g)}$$

(Ratio. 1 : 1 : 2)

(*b*) **Combination between nitrogen and hydrogen:** The two gases lead to the formation of ammonia gas under suitable conditions. The chemical equation is

$$\underset{\text{1 Volume}}{N_2\ (g)} + \underset{\text{3 Volumes}}{3H_2\ (g)} \longrightarrow \underset{\text{2 Volumes}}{2NH_3\ (g)}$$

(Ratio 1 : 3 : 2)

(*v*) **Avogadro's Law:** Avogadro proposed that, equal volumes of gases at the same temperature and pressure should contain equal number of molecules.

For example,
If we consider the reaction of hydrogen and oxygen to produce water, we see that two volumes of hydrogen combine with one volume of oxygen to give two volumes of water without leaving any unreacted oxygen.

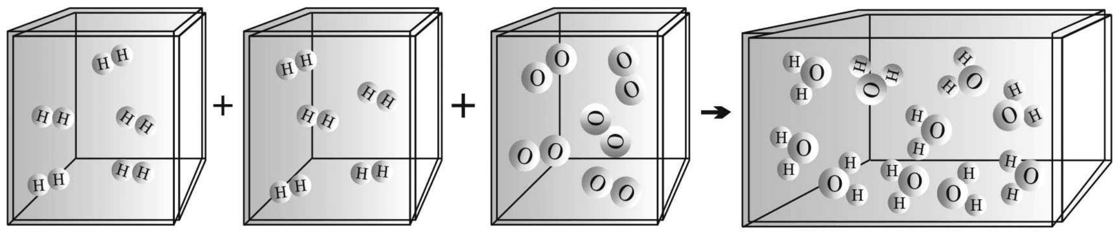


Fig. 1.4 *Two volumes of hydrogen react with one volume of oxygen to give two volumes of water vapour.*

• Dalton's Atomic Theory

In 1808, Dalton published 'A New System of Chemical Philosophy' in which he proposed the following:

1. Matter consists of indivisible atoms.
2. All the atoms of a given element have identical properties including identical mass. Atoms of different elements differ in mass.
3. Compounds are formed when atoms of different elements combine in a fixed ratio.
4. Chemical reactions involve reorganisation of atoms. These are neither created nor destroyed in a chemical reaction.

• Atomic Mass

The atomic mass of an element is the number of times an atom of that element is heavier than an atom of carbon taken as 12. It may be noted that the atomic masses as obtained above are the relative atomic masses and not the actual masses of the atoms.
One atomic mass unit (amu) is equal to 1/12th of the mass of an atom of carbon-12 isotope. It is also known as unified mass.

Average Atomic Mass
Most of the elements exist as isotopes which are different atoms of the same element with different mass numbers and the same atomic number. Therefore, the atomic mass of an element must be its average atomic mass and it may be defined as the average relative mass of an atom of an element as compared to the mass of carbon atoms (C-12) taken as 12*u*.

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 number of its atoms and adding them together.

For example,

Molecular mass of methane (CH_4)
$$= 12.011\ \text{u} + 4\ (1.008\ \text{u})$$
$$= 16.043\ \text{u}$$

Formula Mass

Ionic compounds such as NaCl, KNO_3, Na_2CO_3 etc. do not consist of molecules *i.e.,* single entities but exist as ions closely packed together in a three dimensional space as shown in Fig. 1.5.

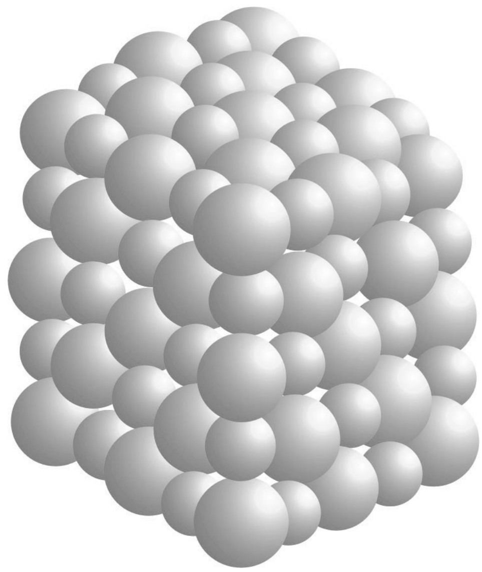


Fig. 1.5 *Packing of Na^+ and Cl^- ions in sodium chloride.*

In such cases, the formula is used to calculate the formula mass instead of molecular mass.
Thus, formula mass of NaCl = Atomic mass of sodium + atomic mass of chlorine
= 23.0 u + 35.5 u = 58.5 u.

• Mole Concept

It is found that one gram atom of any element contains the same number of atoms and one gram molecule of any substance contains the same number of molecules. This number has been experimentally determined and found to be equal to 6.022137×10^{23}

The value is generally called Avogadro's number or Avogadro's constant.

It is usually represented by N_A:

$$\boxed{\text{Avogadro's Number, } N_A = 6.022 \times 10^{23}}$$

• Percentage Composition

One can check the purity of a given sample by analysing this data. Let us understand by taking the example of water (H_2O). Since water contains hydrogen and oxygen, the percentage composition of both these elements can be calculated as follows:

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

$$\text{Molar mass of water} = 18.02 \text{ g}$$

$$\text{Mass \% of hydrogen} = \frac{2 \times 1.008}{18.02} \times 100 = 11.18$$

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

- **Empirical Formula**

The formula of the compound which gives the simplest whole number ratio of the atoms of various elements present in one molecule of the compound.

For example, the formula of hydrogen peroxide is H_2O_2. In order to express its empirical formula, we have to take out a common factor 2. The simplest whole number ratio of the atoms is 1: 1 and the empirical formula is HO. Similarly, the formula of glucose is $C_6H_{12}O_6$. In order to get the simplest whole number of the atoms,

Common factor = 6

The ratio is = 1 : 2 : 1

The empirical formula of glucose = CH_2O

- **Molecular Formula**

The formula of a compound which gives the actual ratio of the atoms of various elements present in one molecule of the compound.

For example, molecular formula of hydrogen peroxide = H_2O_2

and Glucose = $C_6H_{12}O_6$

$$\boxed{\text{Molecular formula} = n \times \text{Empirical formula}}$$

Where n is the common factor and also called multiplying factor. The value of n may be 1, 2, 3, 4, 5, 6 etc.

In case n is 1,

Molecular formula of a compound = Empirical formula of the compound.

- **Stoichiometry and Stoichiometric Calculations**

The word 'stoichiometry' is derived from two Greek words—Stoicheion (meaning element) and metron (meaning measure). Stoichiometry, thus deals with the calculation of masses (sometimes volume also) of the reactants and the products involved in a chemical reaction. Let us consider the combustion of methane. A balanced equation for this reaction is as given below:

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

Here, methane and dioxygen are called reactants and carbon dioxide and water are called products.

(g) indicates that all the reactants and products are gases.

The coefficients 2 for O_2 and H_2O are called stoichiometric coefficients.

Thus, according to the above chemical reaction,

- One mole of CH_4 (g) reacts with two moles of O_2 (g) to give one mole of CO_2 (g) and two moles of H_2O (g).
- One molecule of CH_4 (g) reacts will 2 molecules of O_2 (g) to give one molecule of CO_2 (g) and 2 molecules of H_2O (g).
- 22.4 L of CH_4 (g) reacts with 44.8 L of O_2 (g) to give 22.4 L of CO_2 (g) and 44.82 L of H_2O (g).
- 16 g of CH_4 (g) reacts with 2 × 32 g of O_2 (g) to give 44 g of CO_2 (g) and 2 × 18 g of H_2O (g)

From these relationships, the given data can be interconverted as follows

$$\text{mass} \rightleftharpoons \text{moles} \rightleftharpoons \text{no. of molecules}$$

$$\frac{\text{Mass}}{\text{Volume}} = \text{Density}$$

Limiting Reactant/Reagent

Sometimes, in a chemical equation, the reactants present are not the amount as required according to the balanced equation. The amount of products formed then depends upon the reactant which has reacted completely. This reactant which reacts completely in the reaction is called the limiting reactant or limiting reagent. The reactant which is not consumed completely in the reaction is called excess reactant.

Reactions in Solutions

When the reactions are carried out in solutions, the amount of substance present in its given volume can be expressed in any of the following ways:

1. Mass percent or weight percent (w/w%)
2. Mole fraction
3. Molarity
4. Molality

1. **Mass percent:** It is obtained by using the following relation:

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

2. **Mole fraction:** It is the ratio of number of moles of a particular component to the total number of moles of the solution. For a solution containing n_2 moles of the solute dissolved in n_1 moles of the solvent,
Mole fraction of solute in the solution

$$(x_2) = \frac{n_2}{n_1 + n_2}$$

Mole fraction of solvent in the solution

$$(x_1) = \frac{n_1}{n_1 + n_2}$$

The sum of the mole fractions of the components is equal to 1.
i.e., $x_1 + x_2 = 1$

3. **Molarity:** It is defined as the number of moles of solute in 1 litre of the solution.

$$\text{Thus, Molarity (M)} = \frac{\text{No. of moles of solute}}{\text{Volume of solution in litres}}$$

4. **Molality:** It is defined as the number of moles of solute present in 1 kg of solvent. It is denoted by *m*.

$$\text{Molality } (m) = \frac{\text{No. of moles of solute}}{\text{Mass of the solvent in kg}}$$

Words that Matter

- All substances contain matter which can exist in three states—solid, liquid or gas.
- Matter can also be classified into elements, compounds and mixtures.
- **Element:** An element contains particles of only one type which may be atoms or molecules.
- **Compounds** are formed when atoms of two or more elements combine in a fixed ratio to each other.

- **Mixtures:** Many of the substances present around us are mixtures.
- **Scientific notation:** The measurement of quantities in chemistry are spread over a wide range of 10^{-31} to 10^{23}. Hence, a convenient system of expressing the number in scientific notation is used.
- **Scientific figures:** The uncertainty is taken care of by specifying the number of significant figures in which the observations are reported.
- **Dimensional analysis:** It helps to express the measured quantities in different systems of units.
- **Laws of Chemical Combinations are:**
  (*i*) Law of Conservation of Mass
  (*ii*) Law of Definite Proportions
  (*iii*) Law of Multiple Proportions
  (*iv*) Gay Lussac's Law of Gaseous Volumes
  (*v*) Avogadro's Law.
- **Atomic mass:** The atomic mass of an element is expressed relative to ^{12}C isotope of carbon which has an exact value of 12u.
- **Average atomic mass:** Obtained by taking into account the natural aboundance of different isotopes of that element.
- **Molecular mass:** The molecular mass of a molecule is obtained by taking sum of atomic masses of different atoms present in a molecule.
- **Avogadro number:** The number of atoms, molecules or any other particles present in a given system are expressed in terms of Avogadro constant.
  $$= 6.022 \times 10^{23}$$
- **Balanced chemical equation:** A balanced equation has the same number of atoms of each element on both sides of the equation.
- **Stoichiometry:** The quantitative study of the reactants required or the products formed is called stoichiometry. Using stoichiometric calculations, the amounts of one or more reactants required to produce a particular amount of product can be determined and vice-versa.

NCERT TEXTBOOK QUESTIONS SOLVED

Q1. *Calculate the molecular mass of the following:*
(i) H_2O *(ii)* CO_2 *(iii)* CH_4

Ans. (*i*) Molecular mass of H_2O = 2(1.008 amu) + 16.00 amu
= **18.016 amu**

(*ii*) Molecular mass of CO_2 = 12.01 amu + 2 × 16.00 amu
= **44.01 amu**

(*iii*) Molecular mass of CH_4 = 12.01 amu + 4 (1.008 amu)
= **16.042 amu**

Q2. *Calculate the mass percent of different elements present in sodium sulphate* (Na_2SO_4).

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

Now, Molar mass of Na_2SO_4 = 2 (23.0) + 32.0 + 4 × 16.0
= 142 g mol^{-1}

$$\text{Mass percent of sodium} = \frac{46}{142} \times 100$$
$$= 32.39\ \%$$

$$\text{Mass percent of sulphur} = \frac{32}{142} \times 100$$
$$= 22.54\ \%$$

$$\text{Mass percent of oxygen} = \frac{64}{142} \times 100$$
$$= 45.07\ \%$$

Q3. *Determine the empirical formula of an oxide of Iron which has 69.9 % iron and 30.1 % dioxygen by mass.*

Ans.

Element	*Symbol*	*% by mass*	*Atomic mass*	*Moles of the element (Relative no. of moles)*	*Simplest molar ratio*	*Simplest whole number molar ratio*
Iron	Fe	69.9	55.85	$\frac{69.9}{55.85} = 1.25$	$\frac{1.25}{1.25} = 1$	2
Oxygen	O	30.1	16.00	$\frac{30.1}{16.00} = 1.88$	$\frac{1.88}{1.25} = 1.5$	3

$\therefore$ Empirical formula = $\mathbf{Fe_2O_3}$.

Q4. *Calculate the amount of carbon dioxide that could be produced when*

(i) 1 mole of carbon is burnt in air.

(ii) 1 mole of carbon is burnt in 16 g of dioxygen.

(iii) 2 moles of carbon are burnt in 16 g of dioxygen.

Ans. The balanced equation for the combustion of carbon in dioxygen/air is

$$\underset{\text{1 mole}}{C\,(s)} + \underset{\substack{\text{1 mole} \\ (32\ g)}}{O_2\,(g)} \longrightarrow \underset{\substack{\text{1 mole} \\ (44\ g)}}{CO_2\,(g)}$$

(*i*) In air, combustion is complete. Therefore, CO_2 produced from the combustion of 1 mole of carbon = **44 g.**

(*ii*) As only 16 g of dioxygen is available, it can combine only with 0.5 mole of carbon, *i.e.,* dioxygen is the limiting reactant. Hence, CO_2 produced = **22 g.**

(*iii*) Here again, dioxygen is the limiting reactant. 16 g of dioxygen can combine only with 0.5 mole of carbon. CO_2 produced again is equal to **22 g.**

Q5. *Calculate the mass of sodium acetate (CH_3COONa) required to make 500 mL of 0.375 molar aqueous solution. Molar mass of sodium acetate is 82.0245 g mol^{-1}.*

Ans. 0.375 M aqueous solution means that 1000 mL of the solution contain sodium acetate = 0.375 mole

$\therefore$ 500 mL of the solution should contain sodium acetate = $\frac{0.375}{2}$ mole

Molar mass of sodium acetate = 82.0245 g mol^{-1}

$\therefore$ Mass of sodium acetate required = $\frac{0.375}{2}$ mole × 82.0245 g mol^{-1} = **15.380 g.**

Q6. *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 percent of nitric acid in it is being 69%.*

Ans. Mass percent of 69% means that 100 g of nitric acid solution contain 69 g of nitric acid by mass.

Molar mass of nitric acid HNO_3 = 1 + 14 + 48 = 63 g mol^{-1}

$\therefore$ Moles in 69 g HNO_3 = $\dfrac{69\text{ g}}{63\text{ g mol}^{-1}}$ = 1.095 mole

Volume of 100 g nitric acid solution = $\dfrac{100\text{ g}}{1.41\text{ g mL}^{-1}}$ = 70.92 mL = 0.07092 L

$\therefore$ Conc. of HNO_3 in moles per litre = $\dfrac{1.095\text{ mole}}{0.07092\text{ L}}$ = **15.44 M.**

Q7. *How much copper can be obtained from 100 g of copper sulphate ($CuSO_4$)? (Atomic mass of Cu = 63.5 amu)*

Ans. 1 mole of $CuSO_4$ contains 1 mole (1 g atom) of Cu

Molar mass of $CuSO_4$ = 63.5 + 32 + 4 × 16 = 159.5 g mol^{-1}

Thus, Cu that can be obtained from 159.5 g of $CuSO_4$ = 63.5 g

$\therefore$ Cu that can be obtained from 100 g of $CuSO_4$ = $\dfrac{63.5}{159.5} \times 100$ g = **39.81 g.**

Q8. *Determine the molecular formula of an oxide of iron in which the mass percent of iron and oxygen are 69.9 and 30.1 respectively. Given that the molar mass of the oxide is 159.8 g mol^{-1} (Atomic mass: Fe = 55.85, O = 16.00 amu)*

Calculation of Empirical Formula. See Q3.

Ans. Empirical formula mass of Fe_2O_3 = 2 × 55.85 + 3 × 16.00 = 159.7 g mol^{-1}

$$n = \frac{\text{Molar mass}}{\text{Empirical formula mass}} = \frac{159.8}{159.7} = 1$$

Hence, molecular formula is same as empirical formula, viz., $\mathbf{Fe_2O_3}$.

Q9. *Calculate the atomic mass (average) of chlorine using the following data:*

	% Natural Abundance	*Molar Mass*
^{35}Cl	*75.77*	*34.9689*
^{37}Cl	*24.23*	*36.9659*

Ans. Fractional abundance of ^{35}Cl = 0.7577, Molar mass = 34.9689

Fractional abundance of ^{37}Cl = 0.2423, Molar mass = 36.9659

$\therefore$ Average atomic mass = (0.7577) (34.9689 amu) + (0.2423) (36.9659 amu)

= 26.4959 + 8.9568 = **35.4527**

Q10. *In three moles of ethane (C_2H_6), calculate the following:*

(i) Number of moles of carbon atoms *(ii) Number of moles of hydrogen atoms*

(iii) Number of molecules of ethane

Ans. (*i*) 1 mole of C_2H_6 contains 2 moles of carbon atoms

$\therefore$ 3 moles of C_2H_6 will C-atoms = **6 moles**

(*ii*) 1 mole of C_2H_6 contains 6 moles of hydrogen atoms

$\therefore$ 3 moles of C_2H_6 will contain H-atoms = **18 moles** (*iii*) 1 mole of C_2H_6 contains Avogadro's no., *i.e.,* 6.02×10^{23} molecules

$\therefore$ 3 moles of C_2H_6 will contain ethane molecules = $3 \times 6.02 \times 10^{23}$

= $\mathbf{18.06 \times 10^{23}}$ **molecules**

Q11. *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 2 L?*

Ans. Molar mass of sugar ($C_{12}H_{22}O_{11}$) = $12 \times 12 + 22 \times 1 + 11 \times 16 = 342$ g mol^{-1}

$$\text{No. of moles in 20 g of sugar} = \frac{20\text{ g}}{342\text{g mol}^{-1}} = 0.0585 \text{ mole}$$

$$\text{Molar concentration} = \frac{\text{Moles of solute}}{\text{Volume of sol in L}} = \frac{0.0585}{2\text{ L}} = 0.0293 \text{ mol L}^{-1} = \mathbf{0.0293\ M.}$$

Q12. *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. Molar mass of methanol (CH_3OH) = 32 g mol^{-1} = 0.032 kg mol^{-1}

$$\text{Molarity of the given solution} = \frac{0.793\text{ kg L}^{-1}}{0.032\text{ kg mol}^{-1}} = 24.78 \text{ mol L}^{-1}$$

Applying $M_1 \times V_1 = M_2V_2$

(Given solution) (Solution to be prepared)

$24.78 \times V_1 = 0.25 \times 2.5$ L or $V_1 = 0.02522$ L = **25.22 mL**

Q13. *Pressure is determined as force per unit area of the surface. The S.I. unit of pressure, pascal, is as shown below:*

1 Pa = 1 Nm^{-2}

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

Ans. Pressure is the force (*i.e.,* weight) acting per unit area

But weight = *mg*

$$\therefore \quad \text{Pressure} = \text{Weight per unit area} = \frac{1034\text{ g} \times 9.8\text{ m s}^{-2}}{\text{cm}^2}$$

$$= \frac{1034\text{ g} \times 9.8\text{ ms}^{-2}}{\text{cm}^2} \times \frac{1\text{ kg}}{1000\text{ g}} \times \frac{100\text{ cm}}{1\text{ m}} \times \frac{100\text{ cm}}{1\text{ m}} \times 1 \times \frac{1\text{ N}}{\text{kg ms}^{-2}} \times \frac{1\text{ Pa}}{1\text{ Nm}^{-2}}$$

= $\mathbf{1.01332 \times 10^5}$ **Pa.**

Q14. *What is the S.I. unit of mass?*

Ans. S.I. unit of mass is kilogram (kg).

Q15. *Match the following prefixes with their multiples:*

Prefixes	*Multiples*
(i) micro	10^6
(ii) deca	10^9
(iii) mega	10^{-6}
(iv) giga	10^{-15}
(v) femto	10

Ans. micro = 10^{-6}, deca = 10, mega = 10^6, giga = 10^9, femto = 10^{-15}.

Q16. *What do you mean by significant figures?*

Ans. The digits in a properly recorded measurement are known as significant figures. It is also defined as follows. The total numbers of figures in a number including the last digit whose value is uncertain is called number of significant figures.

Q17. *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 ppm means 15 parts in million (10^6) parts

$$\therefore \quad \% \text{ by mass} = \frac{15}{10^6} \times 100 = 15 \times 10^{-4} = \mathbf{1.5 \times 10^{-3}\ \%}$$

(*ii*) Molar mass of chloroform ($CHCl_3$) = 12 + 1 + 3 × 35.5 = 119.5 g mol^{-1}

100 g of the sample contain chloroform = 1.5×10^{-3} g

$\therefore$ 1000 g (1 kg) of the sample will contain chloroform = 1.5×10^{-2} g

$$= \frac{1.5 \times 10^{-2}}{119.5} = 1.26 \times 10^{-4} \text{ mole}$$

$\therefore$ Molality = $\mathbf{1.266 \times 10^{-4}}$ **m.**

Q18. *Express the following in scientific notation:*

(i) 0.0048	*(ii) 234,000*	*(iii) 8008*	*(iv) 500.0*
(v) 6.0012			

Ans.

(*i*) 4.8×10^{-3}	(*ii*) 2.34×10^5	(*iii*) 8.008×10^3	(*iv*) 5.000×10^2
(*v*) 6.0012×10^0			

Q19. *How many significant figures are present in the following?*

(i) 0.0025	*(ii) 208*	*(iii) 5005*	*(iv) 126,000*
(v) 500.0	*(vi) 2.0034*		

Ans.

(*i*) 2	(*ii*) 3	(*iii*) 4	(*iv*) 3
(*v*) 4	(*vi*) 5.		

Q20. *Round up the following upto three significant figures:*

(i) 34.216	*(ii) 10.4107*	*(iii) 0.04597*	*(iv) 2808*

Ans.

(*i*) 34.2	(*ii*) 10.4	(*iii*) 0.0460	(*iv*) 2810

Q21. *The following data were obtained when dinitrogen and dioxygen react together to form different compounds:*

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

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

(b) Fill in the blanks in the following conversions:

(i) 1 km = ... mm = ... pm *(ii) 1 mg = ... kg = ...ng*

(iii) 1 mL = ... L = ... dm^3

Ans. (*a*) Fixing the mass of dinitrogen as 28 g, masses of dioxygen combined will be 32, 64, 32 and 80 g in the given four oxides. These are in the ratio 1 : 2 : 1 : 5 which is a simple whole number ratio. Hence, the given data obey the **law of multiple proportions.**

(*b*) (*i*) $1 \text{ km} = 1 \text{ km} \times \dfrac{1000 \text{ m}}{1 \text{ km}} \times \dfrac{100 \text{ cm}}{1 \text{ m}} \times \dfrac{10 \text{ mm}}{1 \text{ cm}} = \mathbf{10^6 \text{ mm}}$

$1 \text{ km} = 1 \text{ km} \times \dfrac{1000 \text{ m}}{1 \text{ km}} \times \dfrac{1 \text{ pm}}{10^{-12} \text{ m}} = \mathbf{10^{15} \text{ pm}}$

(*ii*) $1 \text{ mg} = 1 \text{ mg} \times \dfrac{1 \text{ g}}{1000 \text{ mg}} \times \dfrac{1 \text{ kg}}{1000 \text{ g}} = \mathbf{10^{-6} \text{ kg}}$

$1 \text{ mg} = 1 \text{ mg} \times \dfrac{1 \text{ g}}{1000 \text{ mg}} \times \dfrac{1 \text{ ng}}{10^{-9} \text{ g}} = \mathbf{10^{6} \text{ ng}}$

(*iii*) $1 \text{ mL} = 1 \text{ mL} \times \dfrac{1 \text{ L}}{1000 \text{ mL}} = \mathbf{10^{-3} \text{ L}}$

$1 \text{ mL} = 1 \text{ cm}^3 = 1 \text{ cm}^3 \times \dfrac{1 \text{ dm} \times 1 \text{ dm} \times 1 \text{ dm}}{10 \text{ cm} \times 10 \text{ cm} \times 10 \text{ cm}} = 10^{-3} \text{ dm}^3.$

Q22. *If the speed of light is 3.0×10^8 ms^{-1}, calculate the distance covered by light in 2.00 ns.*

Ans. Distance covered = Speed × Time = $3.0 \times 10^8 \text{ ms}^{-1} \times 2.00 \text{ ns}$

$$= 3.0 \times 10^8 \text{ ms}^{-1} \times 2.00 \text{ ns} \times \frac{10^{-9} \text{ s}}{1 \text{ ns}} = 6.00 \times 10^{-1} \text{ m} = \mathbf{0.600 \text{ m}}$$

Q23. *In the reaction, $A + B_2 \longrightarrow AB_2$, identify the limiting reagent, if any, in the following mixtures*

(i) 300 atoms of A + 200 molecules of B
(ii) 2 mol A + 3 mol B
(iii) 100 atoms of A + 100 molecules of B
(iv) 5 mol A + 2.5 mol B
(v) 2.5 mol A + 5 mol B

Ans. (*i*) According to the given reaction, 1 atom of A reacts with 1 molecule of B

∴ 200 molecules of B will react with 200 atoms of A and 100 atoms of A will be left unreacted. Hence, B is the limiting reagent while A is the excess reagent.

(*ii*) According to the given reaction, 1 mol of A reacts with 1 mol of B

∴ 2 mol of A will react with 2 mol of B. Hence, A is the limiting reactant.

(*iii*) No limiting reagent.

(*iv*) 2.5 mol of B will react with 2.5 mol of A. Hence, B is the limiting reagent.

(*v*) 2.5 mol of A will react with 2.5 mol of B. Hence, A is the limiting reagent.

Q24. *Dinitrogen and dihydrogen react with each other to produce ammonia according to the following chemical equation:*

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

(i) Calculate the mass of ammonia produced if 2.00×10^3 g dinitrogen reacts with 1.00×10^3 g dihydrogen
(ii) Will any of the two reactants remain unreacted?
(iii) If yes, which one and what would be its mass?

Ans. (*i*) 1 mol of N_2 *i.e.,* 28 g react with 3 mol of H_2, *i.e.,* 6 g of H_2

$\therefore$ 2000 g of N_2 will react with $H_2 = \frac{6}{28} \times 200$ g = 428.6 g. Thus, N_2 is the limiting reagent while H_2 is the excess reagent.

2 mol of N_2, *i.e.,* 28 g of N_2 produce NH_3 = 2 mol = 34 g

$\therefore$ 2000 g of N_2 will produce $NH_3 = \frac{34}{28} \times 2000$ g = **2428.57 g**

(*ii*) H_2 will remain unreacted.

(*iii*) Mass left unreacted = 1000 g – 428.6 g = 571.4 g

Q25. *How are 0.50 mol Na_2CO_3 and 0.50 M Na_2CO_3 different?*

Ans. Molar mass of $Na_2CO_3 = 2 \times 23 + 12 + 3 \times 16 = 106$g mol^{-1}

0.50 mol Na_2CO_3 means 0.50 × 105 g = 53 g

0.50 M Na_2CO_3 means 0.50 mol, *i.e.,* 53 g Na_2CO_3 are present in 1 litre of the solution.

Q26. *If ten volumes of dihydrogen gas reacts with five volumes of dioxygen gas, how many volumes of water vapour could be produced?*

Ans. H_2 and O_2 react according to the equation

$$2H_2\,(g) + O_2\,(g) \longrightarrow 2\,H_2O\,(g)$$

Thus, 2 volumes of H_2 react with 1 volume of O_2 to produce 2 volumes of water vapour. Hence, 10 volumes of H_2 will react completely with 5 volumes of O_2 to produce **10 volumes** of water vapour.

Q27. *Convert the following into basic units:*

(i) 28.7 pm *(ii) 15.15 μs* *(iii) 25365 mg*

Ans. (*i*) $28.7 \text{ pm} = 28.7 \text{ pm} \times \frac{10^{-12}\text{ m}}{1 \text{ pm}} = \mathbf{2.87 \times 10^{-11}\ m}$

(*ii*) $15.15\ \mu\text{s} = 15.15\ \mu\text{s} \times \frac{10^{-6}\text{ s}}{1\ \mu\text{s}} = \mathbf{1.515 \times 10^{-5}\ s}$

(*iii*) $25365 \text{ mg} = 25365 \text{ mg} \times \frac{1\text{ g}}{1000\text{ mg}} \times \frac{1\text{ kg}}{1000\text{ g}} = \mathbf{2.5365 \times 10^{-2}\ kg}$

Q28. *Which one of the following will have largest number of atoms?*

(i) 1 g Au (s) *(ii) 1 g Na (s)* *(iii) 1 g Li (s)* *(iv) 1 g of Cl_2 (g)*

(Atomic masses: Au = 197, Na = 23, Li = 7, Cl = 35.5 amu)

Ans. (*i*) $1 \text{ g Au} = \frac{1}{197} \text{ mol} = \frac{1}{197} \times 6.02 \times 10^{23}$ atoms

(*ii*) $1 \text{ g Na} = \frac{1}{23} \text{ mol} = \frac{1}{23} \times 6.02 \times 10^{23}$ atoms

(*iii*) $1 \text{ g Li} = \frac{1}{7} \text{ mol} = \frac{1}{7} \times 6.02 \times 10^{23}$ atoms

(*iv*) $1 \text{ g } Cl_2 = \frac{1}{71} \text{ mol} = \frac{1}{71} \times 6.02 \times 10^{23}$ molecules $= \frac{2}{71} \times 6.02 \times 10^{23}$ atoms

Thus, **1 g of Li** has the largest number of atoms.

Q29. *Calculate the molarity of a solution of ethanol in water in which the mole fraction of ethanol is 0.040.*

Ans. $x_{C_2H_5OH} = \dfrac{n\,(C_2H_5OH)}{n\,(C_2H_5OH) + n(H_2O)} = 0.040$ (*Given*) ...(*i*)

The aim is to find number of moles of ethanol in 1 L of the solution which is nearly = 1 L of water (because solution is dilute)

No. of moles in 1 L of water = $\dfrac{1000\text{ g}}{18\text{ g mol}^{-1}}$ = 55.55 moles

Substituting n (H_2O) = 55.55 in eqn (*i*), we get

$$\frac{n\,(C_2H_5OH)}{n\,(C_2H_5OH) + 55.55} = 0.040$$

or 0.96 n (C_2H_5OH) = 55.55 × 0.040 or n (C_2H_5OH) = 2.31 mol

Hence, molarity of the solution = **2.31 M.**

Q30. *What will be the mass of one ^{12}C atom in g?*

Ans. 1 mol of ^{12}C atoms = 6.022×10^{23} atoms = 12 g

Thus, 6.022×10^{23} atoms of ^{12}C have mass = 12g

$\therefore$ 1 atom of ^{12}C will have mass = $\dfrac{12}{6.022 \times 10^{23}}$ g = $\mathbf{1.9927 \times 10^{-23}}$ **g**

Q31. *How many significant figures should be present in the answer of the following calculations?*

(*i*) $\dfrac{0.02856 \times 298.15 \times 0.112}{0.5785}$ (*ii*) *5 × 5.364* (*iii*) *0.0125 + 0.7864 + 0.0215*

Ans. (*i*) The least precise term has 3 significant figures (*i.e.,* in 0.112). Hence, the answer should have 3 significant figures.

(*ii*) Leaving the exact number (5), the second term has 4 significant figures. Hence, the answer should have 4 significant figures.

(*iii*) In the given addition, the least number of decimal places in the term is 4. Hence, the answer should have 4 significant.

Q32. *Use the data given in the following table to calculate the molar mass of naturally occurring argon.*

Isotope	*Isotopic molar mass*	*Abundance*
^{36}Ar	*35.96755 g mol^{-1}*	*0.337*
^{38}Ar	*37.96272 g mol^{-1}*	*0.063*
^{40}Ar	*39.9624 g mol^{-1}*	*99.600*

Ans. Molar mass of Ar = 35.96755 × 0.00337 + 37.96272 × 0.00063 + 39.96924 × 0.99600

= **39.948 g mol^{-1}.**

Q33. *Calculate the number of atoms in each of the following:*

(i) 52 moles of He *(ii) 52 u of He* *(iii) 52 g of He*

Ans. (*i*) 1 mol of He = 6.022×10^{23} atoms

$\therefore$ 52 mol of He = $52 \times 6.022 \times 10^{23}$ atoms = $\mathbf{3.131 \times 10^{25}}$ **atoms**

(*ii*) 1 atom of He = 4 u of He

4 u of He = 1 atom of He

$\therefore$ 52 u of He = $\dfrac{1}{4}$ × 52 atoms = **13 atoms** (*iii*) 1 mole of He = 4 g = 6.022×10^{23} atoms

$$\therefore \quad 52 \text{ g of He} = \frac{6.022 \times 10^{23}}{4} \times 52 \text{ atoms}$$

$$= \mathbf{7.8286 \times 10^{24}} \textbf{ atoms.}$$

Q34. *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 S.T.P.) 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. Amount of carbon in 3.38 g CO_2 = $\frac{12}{44} \times 3.38$ g = 0.9218 g

Amount of hydrogen in 0.690 g H_2O = $\frac{2}{18} \times 0.690$ g = 0.0767 g

As compound contains only C and H, therefore, total mass of the compound
= 0.9218 + 0.0767 g = 0.9985 g

% of C in the compound = $\frac{0.9218}{0.9985} \times 100 = 92.32$

% of H in the cmpound = $\frac{0.0767}{0.9985} \times 100 = 7.68$

Calculation of Empirical Formula

Element	*% by mass*	*Atomic mass*	*Moles of the element*	*Simplest molar ratio*	*Simplest whole no. molar ratio*
C	92.32	12	$\frac{92.32}{12} = 7.69$	1	1
H	7.68	1	$\frac{7.68}{1} = 7.68$	1	1

$\therefore$ Empirical formula = **CH**

10.0 L of the gas at STP weight = 11.6 g

$\therefore$ 22.4 L of the gas at S.T.P will weight = $\frac{11.6}{10.0} \times 22.4 = 25.984 \text{ g} \approx 26$ g

$\therefore$ Molar mass = **26 g mol^{-1}**

Empirical formula mass of CH = 12 + 1 = 13

$\therefore$ $n = \frac{\text{Molecular mass}}{\text{E.F. mass}} = \frac{26}{13} = 2$ $\quad\therefore$ Molecular formula = 2 × CH = $\mathbf{C_2H_2}$

Q35. *Calcium carbonate reacts with aqueous HCl according to the reaction*

$$CaCO_3\,(s) + 2HCl\,(aq) \longrightarrow CaCl_2\,(aq) + CO_2\,(g) + H_2O\,(l).$$

What mass of $CaCO_3$ is required to react completely with 25 mL of 0.75 M HCl?

Ans. Step 1. *To calculate mass of HCl in 25 mL of 0.75 m HCl*

1000 mL of 0.75 M HCl contain HCl = 0.75 mol = 0.75 × 36.5 g = 24.375 g

$\therefore$ 25 mL of 0.75 HCl will contain HCl = $\frac{24.375}{1000} \times 25$ g = 0.6844 g.

Step 2. *To calculate mass of $CaCO_3$ reacting completely with 0.9125 g of HCl*

$$CaCO_3\ (s) + 2HCl\ (aq) \longrightarrow CaCl_2\ (aq) + CO_2\ (g) + H_2O\ (l)$$

2 mol of HCl, *i.e.,* 2 × 36.5 g = 73 g HCl react completely with $CaCO_3$ = 1 mol = 100 g

∴ 0.6844 g HCl will react completely with $CaCO_3 = \frac{100}{73} \times 0.6844$ g = **0.938 g.**

Q36. *Chlorine is prepared in the laboratory by treating manganese dioxide (MnO_2) with aqueous hydrochloric acid according to the reaction.*

$$4\ HCl\ (aq) + MnO_2\ (s) \longrightarrow 2\ H_2O\ (l) + MnCl_2\ (aq) + Cl_2\ (g)$$

How many grams of HCl react with 5.0 g of manganese dioxide? (Atomic mass of Mn = 55 u)

Ans. 1 mole of MnO_2, *i.e.,* 55 + 32 = 87 g MnO_2 react with 4 moles of HCl, *i.e.,* 4 × 36.5 g = 146 g of HCl.

∴ 5.0 g of MnO_2 will react with HCl = $\frac{146}{87} \times 5.0$ g = **8.40 g**

MORE QUESTIONS SOLVED

I. VERY SHORT ANSWER TYPE QUESTIONS

Q1. *What is the SI unit of molarity?*

Ans. SI unit of molarity = mol dm^{-3}

Q2. *What do you understand by stoichiometric coefficients in a chemical equation?*

Ans. The coefficients of reactant and product involved in a chemical equation represented by the balanced form, are known as stoichiometric coefficients.

For example,

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

The stoichiometric coefficients are 1, 3 and 2 respectively.

Q3. *Give an example of a molecule in which the ratio of the molecular formula is six times the empirical formula.*

Ans. The compound is glucose. Its molecular formula is $C_6H_{12}O_6$, while empirical formula is CH_2O.

Q4. *What is an atom according to Dalton's atomic theory?*

Ans. According to Dalton's atomic theory, an atom is the ultimate particle of matter which cannot be further divided.

Q5. *Why air is not always regarded as homogeneous mixture?*

Ans. This is due to the presence of dust particles.

Q6. *Define the term 'unit' of measurement.*

Ans. It is defined as the standard of reference chosen to measure a physical quantity.

Q7. *Define law of conservation of mass.*

Ans. It states that matter can neither be created nor destroyed.

Q8. *How is empirical formula of a compound related to its molecular formula?*

Ans. Molecular formula = (Empirical formula)$_n$

where n is positive integer.

Q9. *How many oxygen atoms are there in 18 g of water?*

Ans. Molar mass of water is 18 g/mol.

Number of oxygen atoms is 18 g of water = 6.02 × 10^{23}

Q10. *Name two factors that introduce uncertainty into measured figures.*

Ans. (*i*) Reliability of measuring instrument.

(*ii*) Skill of the person making the measurement.

Q11. *State Avogadro's law.*

Ans. Equal volumes of all gases under the conditions of same temperature and pressure contain the same number of molecules.

Q12. *How are 0.5 ml of NaOH differents from 0.5 M of NaOH?*

Ans. 0.5 ml of NaOH means 0.5 mole (20.0 g) of NaOH, 0.5M of NaOH means that 0.5 mole (20.0g) of NaOH are dissolved in 1L of its solution.

Q13. *What is one a.m.u. or one 'u'?*

Ans. 1 a.m.u. or 1 u = $\frac{1}{12}$th mass of an atom of carbon 12.

Q14. What is the number of significant figures in 1.050×10^4?

Ans. Four.

II. SHORT ANSWER TYPE QUESTIONS

Q1. *Define molality. How does molality depend on temperature?*

Ans. Molality is defined as the moles of solute per kilogram of solvent.

$$\text{Molality} = m = \frac{\text{Moles of solute}}{\text{Mass of solvent (in kg)}}$$

Molality of a solution does not depend on temperature.

Q2. *Convert 2.6 minutes in seconds.*

Ans. We know that, 1 min = 60 s

$$\text{Conversion factor} = \frac{60\text{ s}}{(1\text{ min})}$$

$$2.6\text{ min} = 2.6\text{ min} \times \text{conversion factor} = 2.6 \times \frac{60\text{ s}}{1\text{ min}} = \mathbf{156\ s.}$$

Q3. *Express the following up to four significant figures.*

(i) 6.5089 *(ii) 32.3928* *(iii) 8.721×10^4* *(iv) 2000*

Ans. (*i*) 6.509 (*ii*) 32.39 (*iii*) 8.721×10^4 (*iv*) 2.000×10^3

Q4. *Calculate the number of moles in each of the following.*

(i) 392 g of sulphuric acid

(ii) 44.8 litres of sulphur dioxide at N.T.P.

(iii) 6.022×10^{22} molecules of oxygen

(iv) 8g of calcium

Ans. (*i*) *392 g of sulphuric acid*

Molar mass of H_2SO_4 = 2 × 1 + 32 + 4 × 16 = 98 g

98 g of sulphuric acid = 1 mol

$$392\text{ g of sulphuric acid} = 1\text{ mol} \times \frac{392\text{ g}}{(98\text{ g})} = \mathbf{4\ mol}$$

(*ii*) *44.8 litres of sulphur dioxide at N.T.P.*

22.4 litres of sulphur dioxide at N.T.P. = 1 mol

$$44.8\text{ litres of sulphur dioxide at N.T.P.} = \frac{1\text{ mol}}{(22.4\text{L})} \times (44.8\text{ L}) = \mathbf{2.0\ mol}$$

(*iii*) *6.022 × 10^{22} molecules of oxygen*

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

6.022 × 10^{22} molecules of oxygen = 1 mol × $\frac{6.022 \times 10^{22}}{6.022 \times 10^{23}}$ = **0.1 mol**

(*iv*) *8g of calcium*

Gram atomic mass of Ca = 40 g

40 g of calcium = 1 mol

8.0 g of calcium = 1 mol × $\frac{(8.0\ g)}{(40\ g)}$ = **0.2 mol.**

Q5. *A compound on analysis was found to contain C = 34.6%, H = 3.85% and O = 61.55%. Calculate the empirical formula.*

Ans. Step I. *Calculation of simplest whole number ratios of the elements.*

Element	*Percentage*	*Atomic Mass*	*Gram atoms (Moles)*	*Atomic ratio (Molar ratio)*	*Simplest whole no. ratio*
C	34.6	12	$\frac{34.6}{12} = 2.88$	$\frac{2.88}{2.88} = 1$	3
H	3.85	1	$\frac{3.85}{1} = 3.85$	$\frac{3.85}{2.88} = 1.337$ or $\frac{4}{3}$	4
O	61.55	16	$\frac{61.55}{16} = 3.85$	$\frac{3.85}{2.88} = 1.337$ or $\frac{4}{3}$	4

The simplest whole number ratios of the different elements are: C : H : O :: 3 : 4 : 4

Step II. *Writing the empirical formula of the compound.*

The empirical formula of the compound = $\mathbf{C_3H_4O_4}$.

Q6. *Calculate:*

(a) Mass of 2.5 gram atoms of magnesium,

(b) Gram atom in 1.4 grams of nitrogen (Atomic mass Mg = 24, N = 14)

Ans. (*a*) 1 gram atom of Mg = 24g

2.5 gram atoms of Mg = 24 × 2.5 = **60g**

(*b*) 1 gram atom of N = 14g;

14g of N = 1 gram atom

1.4g of N = $\frac{1}{14} \times 1.4$ = **0.1 gram atom.**

Q7. *The density of water at room temperature is 1.0g/mL. How many molecules are there in a drop of water if its volume is 0.05 mL?*

Ans. Volume of a drop of water = 0.05 mL

Mass of a drop of water = Volume × density

= (0.05 mL) × (1.0 g/mL) = 0.05 g

Gram molecular mass of water (H_2O) = 2 × 1 + 16 = 18 g

18 g of water = 1 mol

0.05 g of water = $\frac{1\ mol}{(18\ g)} \times (0.05\ g) =$ 0.0028 mol

No. of molecules present

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

0.0028 mole of water contain molecules = $6.022 \times 10^{23} \times 0.0028$ = $\mathbf{1.68 \times 10^{21}}$ **molecules.**

Q8. *What is the molecular mass of a substance each molecule of which contains 9 atoms of carbon, 13 atoms of hydrogen and 2.33×10^{-23} g other component?*

Ans. Mass of 9 atoms of carbon = 9 × 12 amu = 108 u.

Mass of 13 atoms of hydrogen = 13 × 1 amu = 13 u

Mass of 2.33×10^{-23}g of other component = $(1u) \times \dfrac{(2.33 \times 10^{-23}\, g)}{(1.66 \times 10^{-24}\, g)}$ = 14.04 u

Molecular mass of the substance = (108 + 13 + 14.04) u = **135.04 u.**

III. LONG ANSWER TYPE QUESTIONS

Q1. *Calculate no. of carbon and oxygen atoms present in 11.2 litres of CO_2 at N.T.P.*

Ans. **Step I.** Number of CO_2 molecules in 11.2 litres

22.4 litres of CO_2 at N.T.P. = 1 gram mol

11.2 litres of CO_2 at N.T.P. = $\dfrac{(1 \text{ gram mol})}{(22.4 \text{ litres})} \times (11.2 \text{ litres})$ = 0.5 gram mol

Now 1 gram mole of CO_2 contain molecules = 6.022×10^{23}

$\therefore$ 0.5 gram mole of CO_2 contain molecules = $6.022 \times 10^{23} \times 0.5 = 3.011 \times 10^{23}$

Step II. *Number of carbon and oxygen atoms in 3.011×10^{23} molecules of CO_2*

1 molecule of CO_2 contains carbon atoms = 1

$\therefore$ 3.011×10^{23} molecules of CO_2 will contain carbon atoms = $\mathbf{3.011 \times 10^{23}}$

Similarly, 1 molecule of CO_2 contains oxygen atoms = 2

$\therefore$ 3.011×10^{23} molecules of CO_2 will contain oxygen atoms = $2 \times 3.011 \times 10^{23}$

= $\mathbf{6.022 \times 10^{23}}$ **atoms.**

Q2. *$KClO_3$ on heating decomposes to give KCl and O_2. What is the volume of O_2 at N.T.P liberated by 0.1 mole of $KClO_3$?*

Ans. The chemical equation for the decomposition of $KClO_3$ is

$$\underset{2\text{ mol}}{2KClO_3} \xrightarrow{\text{Heat}} 2KCl + \underset{\substack{3\text{ mol} \\ (3 \times 22.4\text{ L} = 67.2\text{L})}}{3O_2}$$

2 moles of $KClO_3$ evolve O_2 at N.T.P. = 67.2 L

1 mole of $KClO_3$ evolve O_2 at N.T.P. = $\dfrac{67.2}{2}$ L

0.1 mole of $KClO_3$ evolve O_2 at N.T.P. = $\dfrac{67.2}{2} \times 0.1$L = **3.36 L**

Q3. *10 mL of a solution of NaCl containing KCl gave on evaporation 0.93 g of the mixed salt which gave 1.865 g of AgCl by reacting with $AgNO_3$ solution. Calculate the quantity of NaCl in 10 mL of the solution.*

Ans. The chemical equation for the reaction is:

$$\underset{}{AgNO_3} + \underset{23 + 35.5\ (= 58.5\ g)}{NaCl} \longrightarrow AgCl + \underset{108 + 35.5\ (= 143.5\ g)}{NaNO_3}$$

Let the mass of NaCl and KCl in the mixture be respectively a g and b g.

$\therefore \quad a + b = 0.93$ (given)

Let us find AgCl formed on reacting NaCl and KCl with $AgNO_3$ solution.

58.5 g of NaCl give AgCl = 143.5 g

$\therefore \quad a$ g of NaCl will give AgCl = $\dfrac{(143.5\ g)}{(58.5\ g)} \times (a\ g)$

Similarly, 74.5 g of KCl give AgCl = 143.5 g

b g of KCl will give AgCl = $\dfrac{(143.5\ g)}{(74.5\ g)} \times (b\ g)$

But mass of AgCl actually formed = 1.865 g (given)

$$\therefore \quad \frac{143.5 \times a}{58.5} + \frac{143.5 \times b}{74.5} = 1.865; \quad \frac{143.5 \times a}{58.5} + \frac{143.5(0.93 - a)}{74.5} = 1.865$$

$$2.453\ a + 1.93(0.93 - a) = 1.865; \ 2.453\ a + 1.795 - 1.93\ a = 1.865$$

$$0.523\ a = 0.07 \quad \text{or} \quad a = \frac{0.07}{0.523} = 0.14$$

Mass of NaCl in the mixture = **0.14 g**

Mass of KCl in the mixture = (0.93 – 0.14) = **0.79 g.**

Q4. *The cost of table salt (NaCl) and table sugar ($C_{12}H_{22}O_{11}$) are ₹ 1 per kg and ₹ 6 per kg respectively. Calculate their cost per mole.*

Ans. (*a*) *Cost of table salt (NaCl) per mole*

Gram molecular mass of NaCl = 23 + 35.5 = 58.5 g

Now, 1000 g of NaCl cost = ₹ 2

$\therefore \quad$ 58.5 g of NaCl will cost = $\dfrac{2}{(1000\ g)} \times (58.5\ g) = 0.117$ Rupee

= 0.117 × 100 = **12 paise (approx.)**

(*b*) *Cost of table sugar ($C_{12}H_{22}O_{11}$) per mole*

Gram molecular mass of ($C_{12}H_{22}O_{11}$) = 12 × 12 + 22 × 1 = 16 × 11

= 144 + 22 + 176 = 342 g

Now, 1000 g of sugar cost = ₹ 6

$\therefore \quad$ 342 g of sugar will cost = $\dfrac{6}{(1000\ g)} \times (342\ g) = 2.052$

= **2.0 Rupees (approx.)**

Q5. *A flask P contains 0.5 mole of oxygen gas. Another flask Q contains 0.4 mole of ozone gas. Which of the two flasks contains greater number of oxygen atoms?*

Ans. 1 molecule of oxygen (O_2) = 2 atoms of oxygen

1 molecule of ozone (O_3) = 3 atoms of oxygen

In flask P: 1 mole of oxygen gas = 6.022×10^{23} molecules

0.5 mole of oxygen gas = $6.022 \times 10^{23} \times 0.5$ molecules

$= 6.022 \times 10^{23} \times 0.5 \times 2$ atoms = $\mathbf{6.022 \times 10^{23}}$ **atoms**

In flask Q: 1 mole of ozone gas = 6.022×10^{23} molecules

0.4 mole of ozone gas = $6.022 \times 10^{23} \times 0.4$ molecules

$= 6.022 \times 10^{23} \times 0.4 \times 3$ atoms = $\mathbf{7.23 \times 10^{22}}$ **atoms**

$\therefore$ Flask Q has a greater number of oxygen atoms as compared to the flask P.

Q6. *Calculate the total number of electrons present in 1.6 g of methane.*

Ans. (*i*) Molar mass of methane (CH_4) = 12 + 4 × 1 = 16 g

16 g of methane contain molecules = 6.022×10^{23}

$$1.6 \text{ g of methane contain molecule} = \frac{6.022 \times 10^{23}}{(16 \text{ g})} \times (1.6 \text{ g}) = 6.022 \times 10^{22}$$

(*ii*) *Number of electrons in 6.022×10^{22} molecules of methane*

1 molecule of methane contains electrons = 6 + 4 = 10

6.022×10^{22} molecules of methane contain electrons

$= 6.022 \times 10^{22} \times 10 = \mathbf{6.022 \times 10^{23}}$.

Q7. *The vapour density of a mixture of NO_2 and N_2O_4 is 38.3 at 27°C. Calculate the number of moles of NO_2 in 100 g of the mixture.*

Ans. Vapour density of the mixture of NO_2 and N_2O_4 = 38.3

Molecular mass of the mixture = 2 × Vapour density = 2 × 38.3 = 76.6 u = 76.6 g

Mass of the mixture = 100 g

$$\text{No. of moles of the mixture} = \frac{100}{76.6}$$

Let the mass of NO_2 in the mixture = x g

$\therefore$ Mass of N_2O_4 in the mixture = $(100 - x)$ g

Molar mass of NO_2 = 14 + 32 = 46 u = 46 g

Molar mass of N_2O_4 = 28 + 64 = 92 u = 92 g

$$\text{No. of moles of } NO_2 = \frac{x}{46}$$

$$\text{No. of moles of } N_2O_4 = \frac{(100 - x)}{92}$$

$$\text{Total no. of moles in the mixture} = \frac{x}{46} + \frac{(100 - x)}{92}$$

$$\text{Equating } (i) \text{ and } (ii), \ \frac{x}{46} + \frac{(100 - x)}{92} = \frac{100}{76.6}$$

$$92x + 46(100 - x) = \frac{100}{76.6} \times 46 \times 92 = 5524.8$$

$$92x - 46x = 5524.8 - 4600 = 924.8.$$

Q8. *The Vapour Density of a gaseous element is 5 times that of oxygen under similar conditions. If the molecule is triatomic, what will be its atomic mass?*

Ans. Molecular mass of oxygen = 32 u

Density of oxygen = $\frac{32}{2} = 16$ u

Density of gaseous element = 16 × 5 = 80 u

Molecular mass of gaseous element = 80 × 2 = 160 u

Atomicity of the element = 3

Atomic mass of the element = $\frac{\text{Molecular Mass}}{\text{Atomicity}} = \frac{160}{3}$ = **53.33 u.**

Q9. *An alloy of iron (53.6%), nickel (45.8%) and manganese (0.6%) has a density of 8.17 g cm^{-3}. Calculate the number of Ni atoms present in the alloy of dimensions 10.0 cm × 20.0 cm × 15.0 cm.*

Ans. *Calculation of mass of nickel (Ni) in the alloy.*

Volume of the alloy = (10.0 cm) × (20.0 cm) × (15.0 cm) = 3000 cm^3

Mass of the alloy = Density × Volume

= (8.17 g cm^{-3}) × (3000 cm^3) = 24510 g

Mass of Ni in the alloy = (24510 g) × $\frac{45.8}{100}$ = **11225.6 g**

Calculation of number of nickel (Ni) atoms in the alloy

Gram atomic mas of Ni = 59 g

59 g of Ni have atoms = 6.022 × 10^{23}

11225.6 g of Ni have atoms = 6.022 × 10^{23} × $\frac{(11.225.6\ \text{g})}{(59.0\ \text{g})}$ = **1.15 × 10^{26} atoms.**

Q10. *A mixture of oxalic acid and formic acid is heated with concentrated H_2SO_4. The gas evolved is collected and on treatment with KOH solution, the volume of solution decreases by 1/6th. Calculate the molar ratios of two acids in the original mixture.*

Ans. Let *x* moles oxalic acid and *y* moles of formic acid be heated with conc. H_2SO_4 according to the following equations

$$\underset{x\ \text{mol}}{\begin{array}{l}\text{COOH}\\ |\\ \text{COOH}\end{array}} \xrightarrow{H_2SO_4/\text{heat}} \underset{x\ \text{mol}}{CO(g)} + \underset{x\ \text{mol}}{CO_2(g)} + H_2O(l)$$

$$\underset{y\ \text{mol}}{HCOOH} \xrightarrow{H_2SO_4/\text{heat}} \underset{y\ \text{mol}}{CO(g)} + H_2O(l)$$

Total moles of gaseous mixture = Moles of CO + Moles of CO_2

= $(x + y)$ mol + x mol = $(2x + y)$ mol.

Now, KOH absorbs only CO_2 *i.e. x* moles and the volume of the solution decreases by

1/6th of its volume. Since equal volumes of gases have equal number of moles according to Avogadro's Law,

$$\therefore \quad \frac{\text{Moles of } CO_2}{\text{Moles of both gases}} = \frac{x}{(2x+y)} = \frac{1}{6}$$

or $6x = 2x + y$ or $4x = y$ or $\frac{y}{x} = 4$

$\therefore$ **Molar ratio of formic acid : oxalic acid = 4 : 1.**

IV. MULTIPLE CHOICE QUESTIONS

1. One mole of CO_2 contains
(*a*) 6.02×10^{23} atoms of *C* (*b*) 3 g of CO_2
(*c*) 6.02×10^{23} atoms of O (*d*) 18.1×10^{23} molecules of CO_2

2. 5.6 litres of oxygen at NTP is equivalent to
(*a*) 1 mole (*b*) $\frac{1}{4}$ mole (*c*) $\frac{1}{8}$ mole (*d*) $\frac{1}{2}$ mole

3. How many grams are contained in 1 gram atom of Na?
(*a*) 13 g (*b*) 1 g (*c*) 23 g (*d*) $\frac{1}{23}$ g

4. 12 g of Mg will react completely with an acid to give:
(*a*) 1 mole of O_2 (*b*) $\frac{1}{2}$ mole of H_2 (*c*) 1 mole of H_2 (*d*) 2 moles of H_2

5. Which of the following has the highest mass?
(*a*) 1 g atom of *C* (*b*) $\frac{1}{2}$ mole of CH_4
(*c*) 10 mL of water (*d*) 3.011×10^{23} atoms of oxygen

6. The empirical formula of sucrose is
(*a*) CH_2O (*b*) CHO (*c*) $C_{12}\,H_{22}\,O_{11}$ (*d*) $C\,(H_2O)_2$

7. The number of significant figures in 0.0101 is
(*a*) 3 (*b*) 2 (*c*) 4 (*d*) 5

8. The number of grams of oxygen in 0.10 mol of $Na_2CO_3 \cdot 10H_2O$ is
(*a*) 20.8 g (*b*) 18 g (*c*) 108 g (*d*) 13 g

9. The mass of an atom of nitrogen is
(*a*) $\frac{14}{6.023 \times 10^{23}}$ (*b*) $\frac{28}{6.023 \times 10^{23}}$ (*c*) $\frac{1}{6.023 \times 10^{23}}$ g (*d*) 14 amu

Ans. 1. (*a*) 2. (*b*) 3. (*c*) 4. (*b*) 5. (*a*)
6. (*c*) 7. (*a*) 8. (*a*) 9. (*b*)

V. HOTS QUESTIONS

Q1. *What is the difference between 160 cm and 160.0 cm?*

Ans. 160 has three *significant figures while 160.0 has four significant figures.*

Q2. *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_2 and H_2O formed depend upon the amount of methane burnt.

Q3. *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.

Elements	Relative no. of moles	Simplest molar ratio	Simplest whole no. molar ratio
A	1.25	$\frac{1.25}{1.25} = 1$	2
B	1.88	$\frac{1.88}{1.25} = 1.5$	3

(*a*) Atomic mass of A = $\frac{70}{1.25} = 56$

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

(*b*) Molecular mass of the compound = 160

Molecular formula of the compound = Fe_2O_3

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