1 Some Basic Concepts of Chemistry

Before going out on a sunny day, you make it a point to wear sunscreen. Even the principle underlying the operation of sunscreen has a chemistry background. The sunscreen filters incoming ultraviolet rays by combining organic and inorganic compounds.

Topic Notes

- *Matter, its Properties and Measurements*
- *Mole Concept, Stoichiometry, and its Calculation*

MATTER, ITS PROPERTIES AND MEASUREMENTS 1

TOPIC 1

IMPORTANCE OF CHEMISTRY

Chemistry is one such disciplinary branch of science that has helped us fulfil all our requirements, be it in the field of the food industry, agriculture, medicine, plastic, dyes, etc.

Applications of chemistry in various fields include:

(1) Study of polymers and their impact on the environment can be done effectively.

(2) In the manufacture of soaps, dyes, fertilizers, and various inorganic (metal-based) and organic (carbon-based) chemicals, thereby contributing to the nation's economy.

(3) In the manufacture of optical fibres, and various large-scale solid state devices.

(4) In the field of medicine, it has found application in the synthesis of drugs that can help cure cancer (for example: Cisplatin, Taxol), AIDS (for example: AZT – Azidothymidine), etc.

(5) Synthesizing ceramics used in automotive and aerospace engineering.

(6) To study the environmental impact of various by-products produced during synthesis procedures and how to render them safe. Thus, alternatives to hazardous chemicals like CFCs and other chain detergents, that are otherwise harmful to the environment, have been successfully synthesized with the help of chemistry.

TOPIC 2

NATURE OF MATTER

Matter is anything that occupies space and has mass and presence of which can be felt by any one or more of our five senses. Everything around us, for example, book, pen, table, iron, gold, plastics, wood, water and air are composed of matter because they have mass and they occupy space.

Matter can be classified in two ways:

Physical Classification of Matter

Depending upon the physical properties, matter can be classified into three states -

(1) Solid state

(2) Liquid state

(3) Gaseous state

Let's compare some differentiating properties of the three states of matter.

Table: Properties of Three States of Matter

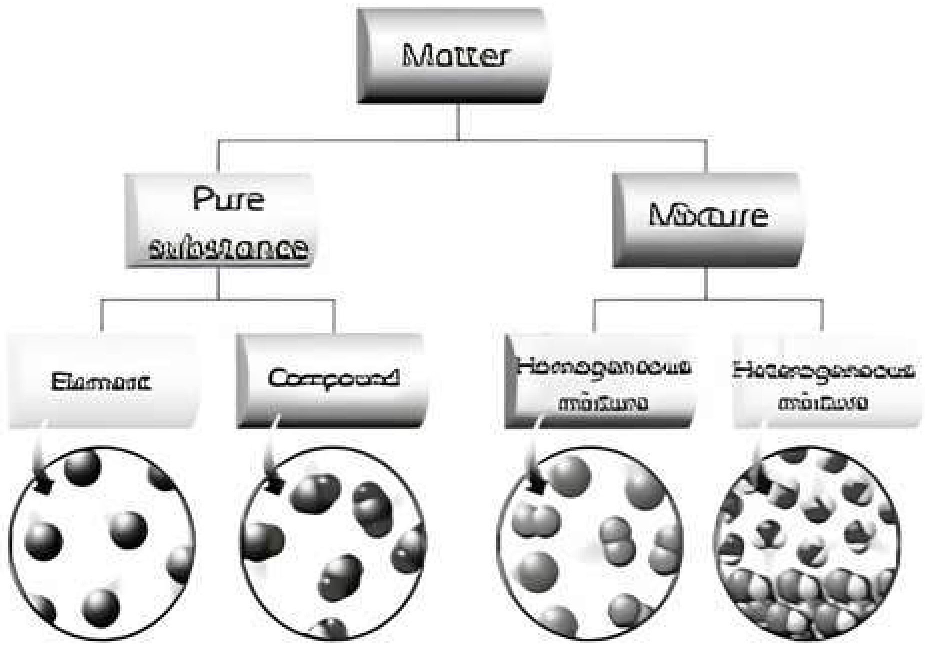
S. No.	Property	Solid	Liquid	Gas
(1)	Intermolecular attraction	Molecules are held together by strong intermolecular forces of attraction.	Attractive forces between molecules are intermediate.	Attractive forces between molecules are very weak.
(2)	Density	Solids have very high density.	Intermediate density to solids and gases.	Possess very low density.
(3)	Fluidity	Does not flow easily.	Flows easily.	Flows easily.
(4)	Diffusion	Very low rate of diffusion.	Moderate rate of diffusion.	Molecules display a very high rate of diffusion.
(5)	Kinetic energy of particles	Particles have very low kinetic energy.	Kinetic energy of particles is intermediate between solid and gas.	Molecules possess very high kinetic energy.

Three states of matter are interconvertible by changing temperature and pressure conditions.

$$\text{Solid} \underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}} \text{Liquid} \underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}} \text{Gas}$$

Chemical Classification of Matter

Based on the chemical composition, matter can be classified as:



Classification of Matter

Mixture

When two or more substances are mixed in any ratio, they form a mixture. A mixture can be further classified as either homogeneous or heterogeneous depending upon the composition of the mixture.

Homogeneous Mixture: If the mixture has a uniform composition throughout, it is referred to as a homogenous mixture. *Eg.*, air, seawater, brass, etc.

Heterogeneous Mixture: If the composition is not uniform throughout the mixture, it is referred to as a heterogeneous mixture. *Eg.*, a mixture of salt and sugar.

Pure Substance

Pure substances consist of a single type of particle. The constituents of pure substances cannot be separated by simple physical methods like filtration, evaporation, distillation, sublimation, mechanical separation, etc. These can only be separated by chemical methods. A pure substance can be classified as an element or a compound. Pure substances are homogeneous since they have a fixed composition.

(1) **Elements:** An element is the simplest form of pure substance which can neither be decomposed into nor built from simpler substances by ordinary physical or chemical methods. It contains only one type of particle, which may be an atom or a molecule.

The elements are further classified into three categories based on their physical and chemical properties.

(i) **Metals:** Metals are those elements that have a lustrous appearance. They are good heat and electrical conductors, as well as malleable and ductile (can be drawn in wire). Some common examples of metals are silver, aluminium, gold, iron, copper, etc.

(ii) **Non-metals:** Non-metals are those elements that do not have shine. They are fragile and have poor heat and electrical conductors (except graphite). *Eg.*, carbon, hydrogen, oxygen, nitrogen, etc.

(iii) **Metalloids:** Metalloids are elements that have properties that are similar to both metals and non-metals. *Eg.*, Bismuth, antimony, arsenic, etc.

(2) **Compounds:** It is formed by atoms of different elements combined in a fixed ratio by mass. They can be either organic compounds or inorganic compounds.

(i) **Organic Compounds:** These are compounds that contain carbon as well as a few additional elements such as hydrogen, nitrogen and sulphur. Plants and animals (living sources) were the first to receive these.

(ii) **Inorganic Compounds:** Those compounds that contain two or more elements out of the currently known 118 elements. Inorganic compounds are obtained from non-living sources such as rocks, minerals, etc.

TOPIC 3

LAWS OF CHEMICAL COMBINATION

Law of Conservation of Mass

It was given by Antoine Lavoisier in 1789, which states that matter can neither be created nor be destroyed. It is just converted from one form to another. Thus, this law is also called the law of indestructibility of matter.

We can also say that:

Sum of masses of reactants = Sum of masses of products.

Example 1.1: 10 grams of calcium carbonate ($CaCO_3$) produces 3.8 grams of carbon dioxide (CO_2) and 6.2 grams of calcium oxide (CaO). Represent this reaction in terms of the law of conservation of mass.

Ans. Given,

Amount of $CaCO_3$ = 10 g

Amount of CO_2 = 3.8 g

Amount of CaO = 6.2 g

According to the law of conservation of mass:
Mass of reactants = Mass of products

$\therefore$ 10 grams of $CaCO_3$ = 3.8 grams of CO_2 + 6.2 grams of CaO

10 grams of reactant = 10 grams of products

Hence, it is proved that the law of conservation of mass is followed by the above reaction.

Law of Definite Proportions

This law was given by Joseph Proust. He stated that a given compound always contains the same proportion of elements by weight. Compounds always contain the same element in the same proportion, whatever the source may be. Thus, this law is also called the Law of definite composition.

Example 1.2: Students of class XI performed a few experiments in class. Vidisha heated 1.375 g of cupric oxide is in a current of hydrogen, the copper formed was 1.098g. In another experiment Shriya dissolved, 1.179g of copper in nitric acid and resulting copper nitrate was converted into cupric oxide by ignition. The weight of cupric oxide formed is 1.476g. Show that these results illustrate the law of definite proportions.

Ans. Here, first, we have to calculate the percentage composition of oxygen in both experiments using the given quantities.

Given that in the first experiment, the mass of cupric oxide is 1.375 g and the weight of copper left is 1.098 g.

So, the mass of oxygen

$$= 1.375\text{ g} - 1.098\text{ g}$$
$$= 0.277\text{ g}$$

Now, we have to calculate the percentage composition of oxygen in copper oxide (CuO).

$$\text{Percentage of O} = \frac{0.277}{1.375} \times 100$$
$$\Rightarrow \quad \%\text{ of O} = 20.15\%$$

In the second experiment, the weight of copper taken is 1.179 g and the mass of copper oxide (CuO) formed is 1.476 g.

Now, we have to calculate the amount of oxygen present.

$$\text{Oxygen present} = 1.476\text{ g} - 1.179\text{ g}$$
$$= 0.297\text{ g}$$

Now, we have to calculate the percentage composition of oxygen in CuO.

$$\text{Percentage of O} = \frac{0.297}{1.476} \times 100$$
$$\Rightarrow \quad \%\text{ of O} = 20.12\%$$

So, we get to know that in both experiments the percentage composition of oxygen is approximately the same. Thus, the law of definite proportions is followed.

Law of Multiple Proportions

This law was proposed by Dalton, according to which, if two elements can combine to form more than one compound, the mass of one element that combines with a fixed mass of another, is in the ratio of small whole numbers. Example: Hydrogen combines with oxygen to form water, and hydrogen peroxide.

$$\underset{2g}{H_2} + \underset{16g}{\frac{1}{2}O_2} \rightarrow \underset{18g}{H_2O}$$

$$\underset{2g}{H_2} + \underset{32g}{O_2} \rightarrow \underset{34g}{H_2O_2}$$

Here, different masses of oxygen combine with a fixed mass of hydrogen, in a simple ratio of 16 : 32 or 1 : 2.

Gay Lussac's Law of Gaseous Volumes

This law was given by Gay Lussac in 1808. According to this law, when gases combine in a chemical reaction, they do so in a simple ratio by volume, provided they are at the same temperature and pressure. For example, in the following reaction, the ratio of volumes of hydrogen, chlorine, and hydrogen chloride is 1 : 1 : 2 (a simple ratio):

$$\underset{1\text{ vol}}{H_{2(g)}} + \underset{1\text{ vol}}{Cl_{2(g)}} \longrightarrow \underset{2\text{ vol}}{2HCl_{(g)}}$$

Avogadro's Law

Equal volumes of all gases contain equal numbers of molecules under the same temperature and pressure conditions. 22.4 L of every gas contains 6.022×10^{23} molecules at STP (Standard Temperature and Pressure).

Important

- *Law of Reciprocal Proportion: In all chemical reactions, substances always react in the ratio of their equivalent masses.*
- *The ratio of the masses of two elements A and B which combine separately with the fixed mass of the third element C is either the same or some simple multiple of the ratio of the masses in which A and B combine directly with each other.*

Example 1.3: Case Based:

The rules of chemical combination are the fundamental principles that interacting atoms and molecules follow. These interactions can involve a wide range of combinations that occur in a variety of ways. When studying chemistry, you must be aware of four general rules that regulate how atoms interact: The law of mass conservation, the law of constant proportions, the law of multiple proportions, and the law of reciprocal proportions are all examples of laws that govern proportions. Pure elements or mixtures of elements termed 'compounds' are rearranged in chemical processes. Atoms' nuclei shift during nuclear reactions. Even though new substances are generated in both circumstances, mass is preserved.

This is the fundamental principle of the law of mass conservation. The law of constant proportions asserts that the constituents in a chemical compound are always present in specified mass quantities. The law of multiple proportions explains how different elements can be combined in different ratios. The same concept holds true for elements and compounds. When two separate components combine with the same quantity of a third element, the ratio in which they do so will be the same or a multiple of the proportion in which they combine with each other, according to the law of reciprocal proportions.

(A) Which of the following is true regarding the law of conservation of mass?
   (a) Sum of masses of reactants = sum of masses of products
   (b) In a chemical compound, the elements are always present in definite proportions by mass.
   (c) When two elements combine with each other to form compounds, the masses of one of the elements which combine with fixed mass of the other, bear a simple whole number ratio to one another
   (d) If two different elements combine separately with a fixed mass of a third element, the ratio of the masses in which they combine are in simple multiple ratios.

(B) What must be held constant when applying Avogadro's law?
   (a) pressure and temperature
   (b) volume and temperature
   (c) moles and temperature
   (d) pressure and volume

(C) The following data are obtained when dinitrogen and dioxygen react together to form different compounds:

S. No.	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

Which law of chemical combination is obeyed by the above experimental data? [NCERT]

(D) Give an example of a law of multiple proportions.

(E) Assertion (A): The balancing of chemical equations is based on the law of conservation of mass.

Reason (R): Total mass of reactants is equal to total mass of products.
   (a) Both (A) and (R) are true and (R) is the correct explanation of (A).
   (b) Both (A) and (R) are true but (R) is not the correct explanation of (A).
   (c) (A) is true but (R) is false.
   (d) (A) is false but (R) is true.

Ans. (A) *(a) Sum of masses of reactants = sum of masses of products*

Explanation: According to the law of conservation of mass, matter can neither be created nor be destroyed. It is just converted from one form to another. Thus, this law is also called the law of indestructibility of matter. We can also say that: Sum of masses of reactants = sum of masses of products

Option (b) is the statement of the law of constant proportion.

Option (c) is the statement of the law of multiple proportions.

Option (d) is the statement of the law of reciprocal proportion.

(B) *(a) pressure and temperature*

Explanation: According to Avogadro's law equal volumes at the same temperature and pressure contain equal numbers of moles of gases.

(C) When 28g of nitrogen combines with dioxygen with mass 32, 64, 32 and 80 g according to different experiment. The ratio corresponds to 1 : 2 : 1 : 5, it's a simple whole number ratio. This illustrates the law of multiple proportions.

(D) The formation of five oxides of nitrogen is an example of the law of multiple proportions as here the elements nitrogen and oxygen combine in different ratios to form different compounds.

(E) *(a) Both (A) and (R) are true and (R) is the correct explanation of (A).*

Explanation: Both assertion and reason are correct since the law of conservation of mass states that total masses of reactants are equal to total masses of products. Also, the balancing of chemical equations is based on this law since balancing involves the same number of atoms of a compound in the reactant and product side.

TOPIC 4

DALTON'S ATOMIC THEORY

Atoms are considered to be the indivisible particles of matter. John Dalton put forward the theory, called the Dalton's Atomic Theory, whose postulates are listed:

(1) Matter comprises of indivisible particles called atoms.

(2) Atoms of the same element have the same mass and size.

(3) Atoms of different elements have varying properties since they differ in mass and size.

(4) Atoms of different elements combine in simple numerical ratios to form compounds. Kinds of atoms are always the same in a given compound.

(5) Atoms are neither created nor destroyed in a chemical reaction.

Modified Atomic Theory

The postulates of Modified Dalton's theory are:

(1) Atom is not the smallest particle of matter. It is not indivisible, rather it is further composed of some subatomic particles.

(2) Atoms of the same element exhibit similar chemical properties because they have the same atomic numbers, even though they have different masses. (isotopes)

(3) Atoms of different elements may have the same mass, but differ in atomic numbers, therefore exhibit different chemical properties. (isobars)

(4) Atoms of one element can be transformed into atoms of another element. (Artificial transmutation)

(5) There are compounds which do not follow the law of constant proportions. They are referred to as the non-stoichiometric compounds.

(6) Atoms are no longer indestructible. In nuclear reactions, atoms of one element may be changed into another. Mass of an atom is related to energy in accordance with Einstein's equation:

$$E = mc^2$$

Where E = energy;

m = mass;

c = velocity of light

Limitations of Dalton's Atomic Theory

Despite of several advantages, Dalton's Atomic Theory was unable to explain the following points:

(1) Failed to explain the law of gaseous volumes.

(2) Could not explain why atoms of different elements exhibit different properties.

(3) Failed to explain the nature of forces that bind the atoms together in a molecule.

(4) Could not explain why atoms combine to form compounds.

(5) Could not differentiate between the ultimate particle of an element or a compound.

TOPIC 4

ATOMIC AND MOLECULAR MASSES

Atomic mass

Carbon – 12 (^{12}C) is the isotope of carbon which is assigned a mass of exactly 12 atomic mass unit, and mass of all other atoms are given relative to it. One atomic mass unit is defined as mass exactly equal to one-twelfth of the mass of one C – 12 atom.

Average Atomic Mass

For elements that exist in isotopic forms, we need to calculate the average atomic mass.

Important

➥ *The atomic masses listed in the periodic table of elements were really the average atomic masses of the elements.*

Average Mass of Carbon

$$= \frac{98.892 \times 12 + 1.108 \times 13.00335 + 2 \times 10^{-10} \times 14.00317}{100}$$

$$= 12.011 \text{ u}$$

Molecular Mass

The sum of atomic masses of elements present in a molecule is its molecular mass. It is obtained by multiplying the atomic mass of each element by the number of atoms and adding them together.

Eg. Molecular mass of CO_2

$= 12 \times 1 + 16 \times 2$

$= 12 + 32$

$= 44$ u

Example 1.4: Calculate the molecular mass of glucose ($C_6H_{12}O_6$) molecule. [NCERT]

Ans. Molecular mass of glucose ($C_6H_{12}O_6$)

= (6 × atomic mass of carbon) + (12 × atomic mass of hydrogen) + (6 × atomic mass of oxygen)

= (6 x 12.011) + (12 x 1.008) + (6 x 16)

= 72.066 u + 12.096 u + 96 u

= 180.162 u

Formula Mass

Formula mass is calculated for ionic compounds which consist of positively charged ions (cations) and negatively charged ions (anions).

Formula mass of an ionic compound = (number of cations × its atomic mass) + (number of anions × its atomic mass)

Formula mass of NaCl = 23 + 35.5 = 58.5 u

Formula mass expressed in grams is called gram formula mass.

OBJECTIVE Type Questions

[1 mark]

Multiple Choice Questions

1. **Calculate the formula weight of calcium hydroxide such that there exist three significant figures in the result:**

(a) 74.096 g (b) 74.1 g
(c) 73.28 g (d) 70 g [Diksha]

Ans. *(b) 74.1 g*

Explanation: Formula weight of compound

= 1 × atomic weight of Ca + 2 × atomic weight of O + 2 × atomic weight of H

= 40.08 amu + 32.00 amu + 2.016 amu

= 74.096

Rounding off these significant figures, the answer becomes 74.1 g.

Caution

- *In the case of multi-step calculation, students used to round off the numbers in between calculations which leads to incorrect or less precise answers. Instead, they should round off the numbers at the end.*

2. **When a temperature was measured on a Fahrenheit scale, it was found to be 200°F. What do you expect its reading on the Celsius scale?**

(a) 40°C (b) 94°C
(c) 93.3°C (d) 30°C

Ans. *(c) 93.3°C*

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

$$200 = \frac{9}{5}(^\circ C) + 32$$

$$^\circ C = 168 \times \frac{5}{9}$$

$$= 93.3$$

Related Theory

- *Fahrenheit scale is commonly used for expressing the body temperature. The normal body temperature on this scale is taken as 98.4°F.*

3. **The mass of 1.5 mL solution with a density 3.12 g/mL in correct significant figures is:**

(a) 4.7 g (b) 4.680 g
(c) 4680×10^{-3} g (d) 46.80 g

Ans. *(a) 4.7 g*

Explanation: Mass of solution

$= V \times d$

$= (1.5\ mL) \times (3.12\ g\ mL^{-1})$

$= 4.7$ g

4. **Which of the following has maximum mass?**

(a) 1.2 g atom of oxygen
(b) 5.2 g atom of iodine
(c) 5.6 g atom of chlorine
(d) 5.6 g atom of sodium

Ans. *(b) 5.2 g atom of iodine*

Explanation: Since, mass of 1.2 gram atom of oxygen = 1.2 × atomic mass of oxygen

= 1.2 × 16 g

= 19.2 g

Mass of 5.2 gram atom of iodine

= 5.2 × atomic mass of iodine

= 5.2 × 127

= 660.4 g

Molar mass of 5.6 g chlorine

= 5.6 × 35.5 g

= 198.8 g

Molar mass of 2.56 g atom of sodium

= 2.56 × 23 g

= 58.88 g

5. **If a student wants to verify the law of conservation of mass, he/she needs to weigh:**

(a) reactants are only required
(b) products are only required
(c) reactants and the product are a must
(d) contents are not essential

Ans. *(c) reactants and the product are a must*

Explanation: To verify the law of conservation of mass we can't weigh reactants only or products only. Instead, we need to weigh both the reactants and products before and after the reaction.

Related Theory

- *According to the law of conservation of mass, mass can neither be created nor it can be destroyed. It can only be transformed from one form to another.*

6. **After adding the given digits 29.4406, 3.2 and 2.25, the answer will have significant figures:**

(a) three (b) four
(c) two (d) five

Ans. *(a) three*

Explanation: 29.4406 + 3.2 + 2.25 = 34.8906

As 3.2 has the least number of decimal places i.e., one, therefore sum should be reported to one decimal place only. After rounding of the reported sum is 34.9, which has three significant figures.

Assertion-Reason (A-R)

In the following question no. (7-9), a statement of assertion followed by a statement of reason is given. Choose the correct answer out of the following choices:

(a) Both (A) and (R) are true and (R) is the correct explanation of (A).
(b) Both (A) and (R) are true but (R) is not the correct explanation of (A).
(c) (A) is true but (R) is false.
(d) (A) is false but (R) is true.

7. Assertion (A): 22 carat gold is a mixture.
Reason (R): A compound has a fixed composition of the elements present in it.

Ans. *(b) Both (A) and (R) are true but (R) is not the correct explanation of (A).*

Explanation: 22 carat gold is a homogeneous mixture in which gold is mixed with a small amount of copper or silver. So, the assertion is correct. But it is not a compound since it is not formed by combination of elements in a fixed ratio by mass.

8. Assertion (A): The standard unit of expressing mass of atoms is called 'u'.
Reason (R): 'u' represents unit mass.

Ans. *(c) (A) is true but (R) is false.*

Explanation: The standard unit of expressing mass of atoms is called 'u' where u represents unified mass and not unit mass.

Related Theory

➥ *Initially, the term Dalton was used to express the mass of a single atom of carbon. Later on, it was replaced by the term atomic mass unit (amu). At present, the term unified mass 'u' is commonly used to represent it.*

9. Assertion (A): Components of a homogeneous mixture cannot be separated by using physical methods.
Reason (R): Composition of a homogeneous mixture is uniform throughout as the components are mixed uniformly.

Ans. *(d) (A) is false but (R) is true.*

Explanation: The homogeneous mixture is defined as the mixture that has the same proportions of its components throughout the solution in any given sample. It is obtained by physically mixing of the components and thus can be separated by physical methods.

CASE BASED Questions (CBQs)

[4 & 5 marks]

Read the following passages and answer the questions that follow:

10. *Recent studies have revealed that the simplest form of matter is atoms and elements may also be defined as the pure substance which is made of one kind of atoms. Examples are carbon, sulphur, hydrogen, oxygen, etc. A compound is also a pure substance like element but it is made up of two or more elements. For example, in sodium chloride the two elements sodium and chlorine are present in the ratio of 23 : 35.5 by mass. Both elements and compounds are pure substances. But on mixing two or more substances in any ratio, mixture results. For example, air is a mixture of different gasses like nitrogen, oxygen, carbon dioxide, water vapour, etc. Further, mixtures are divided into two categories: homogeneous and heterogeneous.*

(A) Classify the following as pure substances or mixtures: Graphite and Iodized table salt.
(B) Why is tap water considered a mixture while distilled water as a compound?
(C) Why is the gaseous state of ammonia regarded as gas while that of water as vapours?

Ans. (A) Graphite- Pure substance (Element).
Iodized table salt- Mixture(Heterogeneous)

(B) Tap water constitutes some impurities such as dust particles which are normally mixed with it and not combined chemically. In tap water, the constituents are not present in a fixed ratio and hence, it is a mixture. Distilled water contains only water molecules since it is free from impurities, it is therefore, considered as a compound.

(C) Only the gaseous states of those substances are regarded as vapours which are liquid at room temperature. Since ammonia exists as a gas at room temperature. Hence, its gaseous state is called gas while water is a liquid at room temperature. Hence, its gaseous state is called vapours.

11. *Atoms and molecules are so small in size that it is neither possible to count them individually nor possible to determine their mass. These are counted collectively in terms of Avogadro's number. The mass of Avogadro's number of atoms and molecules is known as gram atomic mass and gram molecular mass respectively. The volume occupied by Avogadro's number of molecules of a gas or vapour is known as molar volume.*

(A) Mass of CO_2 is 88 g. The number of atoms of oxygen present in it, is:

(a) 2.41×10^{24}

(b) 1.2×10^{23}

(c) 1.4×10^{23}

(d) 2.41×10^{23}

(B) Calculate the molecular mass of cane sugar.

(a) 350 g (b) 361 g

(c) 342 g (d) 345 g

(C) What will be the number of molecules in one mole of a gas at 100°C and 500 mm pressure?

(a) Less than Avogadro's number

(b) Equal to Avogadro's number

(c) Greater than Avogadro's number

(d) With the change in temperature and pressure, the number of particles will change.

(D) If N_A is Avogadro's number, then the number of valence electrons in 70 g of nitride ions (N^{3-}) is:

(a) 42 N_A (b) 40 N_A

(c) 16 N_A (d) 45 N_A

(E) Choose the correct mass (in grams) of 11.2 L of N_2 at STP.

(a) 13 g (b) 14.5 g

(c) 14 g (d) 15 g

Ans. (A) *(a) 2.41×10^{24}*

Explanation: 44.0 g of CO_2 contains oxygen atoms

$$= 2 \times 6.022 \times 10^{23}$$

Now, 88.0 g of CO_2 contains oxygen atoms

$$= 2 \times 2 \times 6.022 \times 10^{23}$$

$$= 2.41 \times 10^{24} \text{ atoms}$$

(B) *(c) 342 g*

Explanation: Molecular mass of cane sugar $C_{12}H_{22}O_{11}$.

$$= (12 \times 12) + (22 \times 1) + (16 \times 11)$$

$$= 342 \text{ g}$$

(C) *(b) Equal to Avogadro's number*

Explanation: The number of molecules in one mole of a gas will be equal to Avogadro's number. However, any change in temperature and pressure will have no influence on the number of particles present.

(D) *(b) 40 N_A*

Explanation: Moles of N^{3-} ion

$$= \frac{70}{14}$$

$$= 5 \text{ mol}$$

No. of N^{3-} ions = 5 × N_A ions

No. of valence electrons in one

N^{3-} ion = 5 + 3 = 8

Total no. of electrons

$$= 5 \times 8 \times N_A$$

$$= 40\ N_A$$

(E) *(c) 14 g*

Explanation: 22.4 L of N_2 at STP weighs

$$= 28.0 \text{ g}$$

11.2 L of N_2 at STP weighs

$$= \frac{28}{22.4} \times 11.2$$

$$= 14.0 \text{ gm}$$

VERY SHORT ANSWER Type Questions (VSA)

[1 mark]

12. What is the mass in grams of nitric acid molecule HNO_3? **[Diksha]**

Ans. Molar mass of HNO_3

$= 1 \times$ atomic mass of H + 1 × atomic mass of N + 3 × atomic mass of O

$= 1 \times 1 + 1 \times 14.01 + 3 \times 16$

$= 63.01$ g

13. Calculate the mass of 1 L of Hg in kg if its density is 13.6 g/cm^3.

Ans. Mass = volume × density

$= 1000\ cm^3 \times 13.6\ g\ cm^{-3}$

$= 13600$ g

$= 13.6$ kg

14. Calculate the mass of 2.5g atoms of magnesium.

Ans. 1 g atom of Mg = 24 g

2.5 g atoms of Mg = 24 × 2.5

= 60 g

15. Why are atomic masses the average values?

Ans. Most of the elements exist in different isotopic forms. Chlorine has 2 isotopes with mass numbers 35 and 37 existing in the ratio 3 : 1. Hence, the average value is taken.

SHORT ANSWER Type-I Questions (SA-I)

[2 marks]

16. A student performs a titration with different burettes and finds titer values of 25.2 mL, 25.25 mL and 25.0 mL. What is the average titer value based upon the rounding off principle?

Ans. Average titer value

$$= \frac{(25.2 + 25.25 + 25.0)}{3} \text{ mL}$$

$= 25.15$ mL

The maximum number of significant figures to be reported is three. This means that the last digit 5 has to be dropped and the correct answer is 25.2.

17. When 4.2 g $NaHCO_3$ is added to a solution of CH_3COOH (acetic acid) weighing 10g, it is observed that 2.2 g of CO_2 is released into the atmosphere. The residue is found to weigh 12.0 g. Show that these observations are in agreement with the law of conservation of mass.

Ans. The chemical equation for the reaction is:

$NaHCO_3 + CH_3COOH \rightarrow CH_3COONa + H_2O + CO_2$

Mass of reactants = mass of $NaHCO_3$ + mass of CH_3COOH

= 4.2 + 10.0

= 14.2 g

Mass of products = mass of residue + mass of CO_2

= 12.0 + 2.2

= 14.2 g

Since there is no change in mass during the chemical reaction, these observations are in agreement with the law of conservation of mass.

18. 10 mL of H_2 contain 2000 molecules under certain temperature and pressure. Calculate the number of molecules of oxygen whose volume is 150 mL at the same temperature and pressure.

Ans. According to Avogadro's law

10 mL of O_2 also contain = 2000 molecules

150 mL of O_2 will contain

$$= \frac{150 \text{ mL}}{10 \text{ mL}} \times 2000$$

$= 30000$

19. If a piece of copper metal weighs 6.342 g, then what will be its volume? (density of copper 7.6 g/cm^3)

Ans. Mass of copper metal = 6.342 g

Density of copper metal = 7.6 g/cm^3

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

$$\frac{6.342 g}{7.6 g/cm^3} = 0.834\ cm^3$$

The final result has to be reported up to significant figures because the least precise number (7.6) has two significant figures.

∴ In the final result, the digit 4 is dropped and the correct answer = 0.83 cm^3.

20. How many Na atoms are present in 143g of washing soda ($Na_2CO_3.10H_2O$)?

Ans. Molar mass of washing soda ($Na_2CO_3.10H_2O$)

$= 2 \times 23 + 12 + 3 \times 16 + 10 \times 18$

$= 46 + 12 + 48 + 180$

$= 286$ g

286g of washing soda = 1 gram mol

143g of washing soda = $\frac{1}{286} \times 143$

= 0.5 gram mol.

No. of sodium atoms present:

1 gram mole of washing soda contains Na atoms

$= 2 \times 6.022 \times 10^{23}$

0.5 gram mole of washing soda contains Na atoms

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

$= 6.022 \times 10^{23}$

SHORT ANSWER Type-II Questions (SA-II)

[3 marks]

21. Calculate average atomic mass of hydrogen using the following data:

Isotope	Natural Abundance%	Molar mass
^1H	99.985	1
^2H	0.015	2

[NCERT Exemplar]

Ans. Average atomic mass

$$= \frac{(\text{Natural abundance of } ^1H \times \text{atomic mass}) + (\text{Natural abundance of } ^2H \times \text{atomic mass of } ^2H)}{100}$$

$$= \frac{99.985 \times 1 + 0.015 \times 2}{100}$$

$$= \frac{99.985 + 0.030}{100}$$

$$= \frac{100.015}{100}$$

$= 1.00015$ u

22. Chlorine has two isotopes of atomic mass units 34.97 u and 36.97 u. The relative abundances of these two isotopes are 0.735 and 0.245 respectively. Find out the average atomic mass of chlorine.

Ans. Atomic mass of first isotope = 34.97 u

Relative abundance of first isotope = 0.735

Atomic mass of second isotope = 36.97 u

Relative abundance of second isotope = 0.2445

Average atomic mass of chlorine

$$= \frac{(34.97\,u) \times 0.735 + (36.97\,u) \times 0.245}{0.735 + 0.245}$$

$$= \frac{(25.703 + 9.058)u}{0.980}$$

$= 35.47$ u

23. Calculate the following:

(A) Number of gram molecules in 4.9 g of H_2SO_4.

(B) Mass of 0.72 g molecules of CO_2.

Ans. (A) Molecular mass of H_2SO_4

= (2 × Atomic mass of H) + (Atomic mass of S) + (4 × Atomic mass of O)

$= (2 \times 1) + 32 + (4 \times 16)$

= 98 u

Gram molecular mass of H_2SO_4 = 98g

98 g of H_2SO_4 = 1 gram molecule

4.9 g of H_2SO_4 = $\frac{1}{98} \times 4.9$

= 0.05 gram

(B) Molecular mass of CO_2

= Atomic mass of C + 2 × Atomic mass of O

$= 12 + 2 \times 16$

= 44 u

Gram molecular mass of CO_2 = 44 g

1 gram molecule of CO_2 has mass = 44 g

0.72 gram molecule of CO_2 has mass

$= 44 \times 0.72$

= 31.68 g

24. The mass of copper oxide obtained by heating 2.16 g of metallic copper with nitric acid and subsequent ignition was found to be 2.7g. Show that the data illustrates the law of constant composition.

Ans. In the experiment:

Weight of copper = 2.16 g

Weight of copper oxide = 2.70 g

Weight of oxygen = 2.70 – 2.16

= 0.54 g

Percentage of copper

$$= \frac{\text{Weight of copper}}{\text{Weight of copper oxide}} \times 100$$

$$= \frac{2.16 \text{ g}}{2.70 \text{ g}} \times 100$$

= 80%

Percentage of oxygen

$$= \frac{\text{Weight of oxygen}}{\text{Weight of copper oxide}} \times 100$$

$$= \frac{0.54 \text{ g}}{2.70 \text{ g}} \times 100$$

= 20%

Since the ratio by weights of copper and oxygen in the compounds remains the same, the law of definite composition is illustrated.

LONG ANSWER Type Questions (LA)

[**4** & **5** marks]

25. Answer the following:

(A) What is an atom according to Dalton's atomic theory?

(B) All atoms of an element always have the same mass. Do you agree with the above view?

(C) Do atoms always combine in simple whole number ratios to form molecules?

Ans. (A) According to Dalton's atomic theory, an atom is the ultimate particle of matter and cannot be further divided into anything simpler than itself.

(B) No, all the atoms of a particular element do not always have the same mass. For example, carbon has two types of atoms with atomic mass 12 amu and 14 amu (amu or atomic mass unit is a scale for expressing atomic mass).

(C) It is not necessary that the atoms may always combine in a simple whole number ratio to form molecules. For example, the formula of sugar (or sucrose) which we daily consume is $C_{12}H_{22}O_{11}$. The elements C, H and O are present in the ratio 12 : 22 : 11 respectively.

26. Define the law of multiple proportions. Explain it with two examples. **[NCERT Exemplar]**

Ans. The Law of Multiple Proportions states:

"When two elements combine to form two or more than two chemical compounds then the weights of one of elements which combine with a fixed weight of the other, bear a simple ratio to one another."

Examples: Compound of Carbon and Oxygen: C and O combine to form two compounds CO and CO_2.

In CO 12 parts of wt. of C combined with 16 parts by wt. O.

In CO_2, 12 parts of wt. of C combined with 32 parts by wt. of O.

If the weight of C is fixed at 12 parts by weight, then the ratio in the weights of oxygen which combine with the fixed weight of C (= 12) is 16 : 32 or 1 : 2.

Thus, the weight of oxygen bears a simple ratio of 1 : 2 to each other.

Compounds of Sulphur (S) and Oxygen (O):

S forms two oxides with O, viz., SO_2 and SO_3

In SO_2, 32 parts of wt. of S combine with 32 parts by wt. of O.

In SO_3, 32 parts of wt. of S combine with 48 parts by wt. of O.

If the wt. of S is fixed at 32 parts, then the ratio in the weights of oxygen which combine with the fixed wt. of S is 32 : 48 or 2 : 3.

Thus, the weights of oxygen have a simple ratio of 2 : 3 to each other.

27. (A) How will you differentiate between homogeneous and heterogeneous mixtures?

(B) How will you justify that water is a compound and not a mixture?

Ans. (A)

S. No.	Homogeneous mixture	Heterogeneous mixture
(1)	Substances are uniformly distributed.	These substances are unevenly distributed.
(2)	These are not visible to the naked eye but visible through the microscope.	These are easily visible to the naked eye and also through a microscope.
(3)	The particles appear smaller in size.	The particles are either smaller or larger in size.

(4)	These are pure substances.	These are not pure substances.
(5)	They represent the same physical properties.	They do not possess the same physical properties.
(6)	Examples include milk, gasoline, sugar solution, etc.	Examples are a mixture of mud and tap water, etc.

(B) Water is considered to be a compound due to the following reasons:

Water cannot be separated into its constituents hydrogen and oxygen by physical methods.

Properties of water are entirely different from its constituent's hydrogen and oxygen. Hydrogen is neither combustible nor a supporter of combustion while oxygen supports combustion. Water is quite different from the two and it extinguishes fire.

Heat and light are given out when water is formed by burning hydrogen and oxygen.

The composition of water is fixed. Its constituent's hydrogen (H) and oxygen (O) are present in the ratio of 1 : 8 by mass.

Water has a fixed boiling point of 100°C under standard atmospheric pressure of one atmosphere.

MOLE CONCEPT, STOICHIOMETRY, AND ITS CALCULATION

2

TOPIC 1

MOLE CONCEPT

The idea of the mole concept is used to count microscopic entities i.e. (atoms, molecules, electrons, ions, etc.).

Mole (symbol – *mol*) is the SI unit of the amount of substance. One mole represents 6.023×10^{23} particles (atoms, molecules, or ions) and is called the Avogadro Constant or Avogadro number (N_A).

For any substance, one mole denotes the same number of entities. On this basis, chemists selected a standard number, which is equal to the number of atoms present in exactly 12.0 g of carbon (^{12}C).

Thus, one mole is the amount of substance that contains the same number of entities as there are atoms present at 12 g of carbon –12 isotopes.

Mathematically, the mass of ^{12}C atom

$$= 1.992648 \times 10^{-23} \text{ g}$$

1 mole of carbon weighs = 12 g

Hence, the number of 12 C atoms

$$= 6.022 \times 10^{23} \text{ atoms/mol.}$$

So, 1 mole = 6.022×10^{23} atoms/molecules/ions.

Mole and Gram Atomic Mass

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

One mole of 6.022×10^{23} atoms = Gram atomic mass of the element

Mole and Gram Molecular Mass

Mass of 6.022×10^{23} molecules (or one mole molecules) of any substance in grams is equal to its gram molecular mass or one gram molecule.

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

= Gram molecular mass

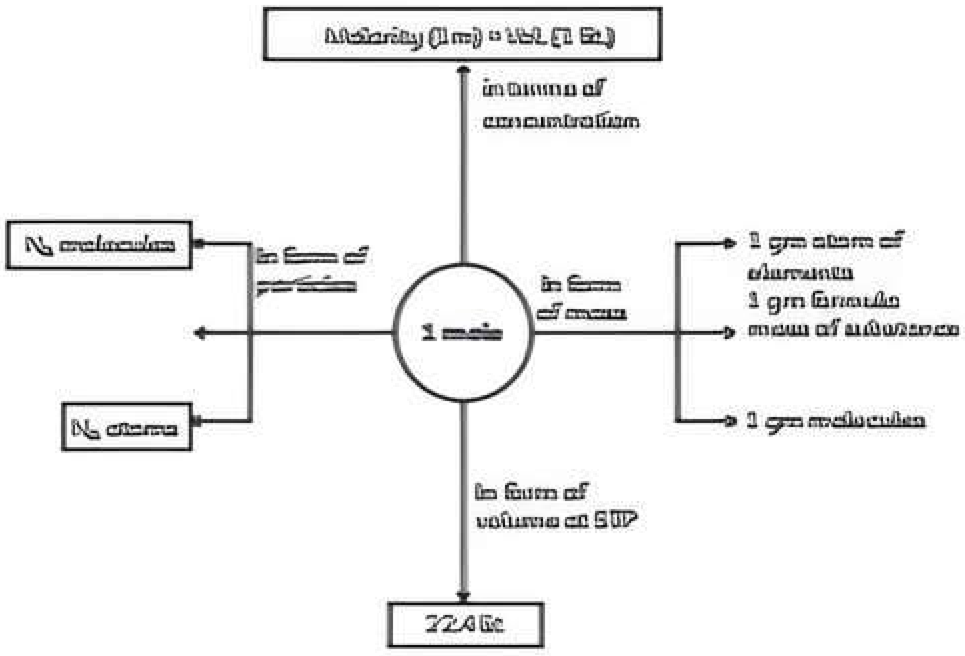
Therefore, we can say, 1 mole of water molecules

= 6.022×10^{23} water molecules

= 18 g.

1 mole of oxygen molecules

= 6.022×10^{23} molecules = 32 g



Relationship between Mole, number of Particles and Mass

Example 2.1: **In the three moles of ethane, calculate the following:**

(A) Number of moles of carbon atoms

(B) Number of moles of hydrogen atoms

(C) Number of molecules of ethane **[NCERT]**

Ans. **(A)** One mole (C_2H_6) has

= 2 two moles of carbon atoms.

Hence, 3 moles of C_2H_6 will have

= 2 × 3

= 6 moles

(B) One mole of C_2H_6 has

= 6 moles of hydrogen atoms.

Hence, 3 moles of C_2H_6 will have

= 3 × 6

= 18 moles

(C) One mole of C_2H_6 has

= 6.022×10^{23} molecules of ethane.

Hence, 3 moles of C_2H_6 will have

= $3 \times 6.023 \times 10^{23}$ molecules

= 18.069×10^{23} molecules

Moles in Case of Ionic Compounds

For ionic compounds formula mass is used instead of using molecular mass. The mass of one mole formula unit in grams is equal to formula mass expressed in grams or gram formula mass of the compound.

Thus, mass of 6.022×10^{23} formula units (or one mole formula units) of any ionic substance in grams is equal to its gram formula mass.

Eg. a mole of NaCl equals to 58.5 g (one gram formula mass) and contains 6.022×10^{23} formula units of NaCl or 6.022×10^{23} Na^+ ions and 6.023×10^{23} Cl^- ions.

Moles in case of Gases

In the case of gases, a mole is defined as the amount of gas that has a volume of 22.4 L at STP, also known as its molar volume.

Eg. 1 mole of Oxygen gas = 22.4 Litres of Oxygen at STP = 32 g

1 mole of Carbon dioxide gas = 22.4 Litres of CO_2 at STP = 44 g

Molar Mass

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

The molar mass of a substance is the mass of one mole in grams. Its units are *g mol^{-1}* and *kg mol^{-1}*. The atomic/molecular/formula mass is quantitatively equal to the molar mass in grams. Molar mass (in grams) is equal to atomic/ molecular/ formula mass in (u) unified mass. A compound's molar mass can be calculated by summing the atomic masses of all the atoms in its molecule.

Eg. molar mass of water = 18.02 g mol^{-1}.

The molar mass of sodium chloride = 58.5 g mol^{-1}.

Important

➥ *The molar mass of a substance is the mass of one mole in grams. Its units are g mol^{-1} and kg mol^{-1}. The atomic/ molecular/formula mass is quantitatively equal to the molar mass in grams.*

Eg. molar mass of water = 18.02 g mol^{-1}.

The following relationships can be summarised as follows:

The number of atoms in a mole

$$= \frac{\text{mass of an element}}{\text{atomic mass}}$$

Number of moles of molecule

$$= \frac{\text{mass of molecule}}{\text{molecular mass}}$$

$$\text{Number of moles of gas} = \frac{\text{volume of the gas (STP)}}{22.4\text{ L}}$$

$$1 \text{ mole} = 6.022 \times 10^{23} \frac{\text{atoms}}{\text{molecules}}$$

$$= \frac{\text{Gram atomic mass}}{\text{molecular mass}}$$

Gram atomic mass / molecular mass

Number of molecules = Number of moles $\times N_A$

where, N_A = Avogadro number

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

Number of atoms = Number of molecules × atomicity (or number of atoms in the molecular formula or in 1 mole).

Example 2.2: How much Copper can be obtained from 100 g of Copper Sulphate ($CuSO_4$)? [NCERT]

Ans. The molar mass of $CuSO_4$ is obtained as:

= Atomic mass of Cu + Atomic mass of S + Cu + 4 × Atomic mass of O.

$= 63.5 + 32 + 4 \times 16$

$= 159.5$ g mol^{-1}

The amount of Cu present in 159.5 g of $CuSO_4$

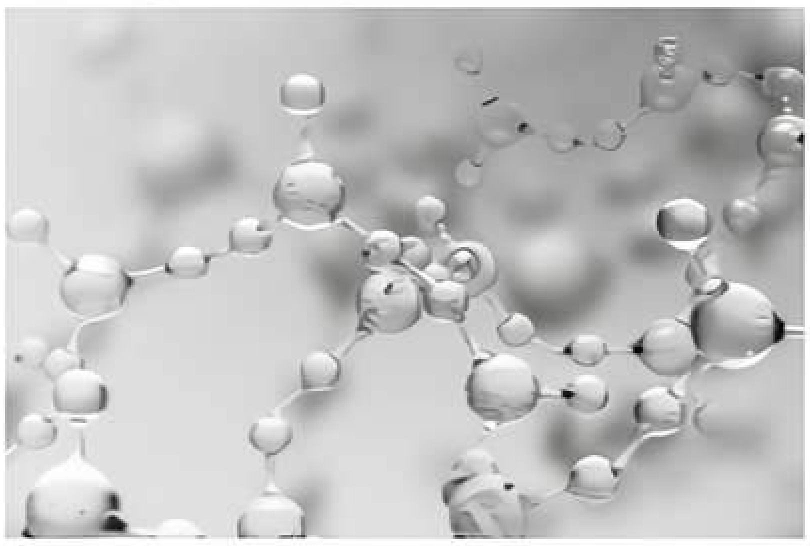
$= 63.5$ g.

Hence, the amount of Cu in 100 g of $CuSO_4$,

$$= \frac{63.5}{159.5} \times 100 = 39.8 \text{ g}$$

Example 2.3: Case Based:

Chemistry is the study of how atoms and molecules interact with one another at the atomic level. The mole idea, which we will discuss here, bridges this gap by linking the mass of a single atom or molecule in amu to the mass of a huge collection of similar molecules in grams. A substance's molecular mass is the sum of the average masses of the atoms in one molecule of that substance. It is computed by multiplying the atomic masses of the elements in the material by their subscripts (written or inferred) in the molecular formula. Any reasonably quantifiable mass of an element or compound comprises an unusually large number of atoms, molecules, or ions, necessitating the use of an unusually big numerical unit to count them. This is accomplished by the use of a mole. A mole is defined as the amount of a substance that contains the number of carbon atoms in exactly 12 g of isotropic pure carbon-12. According to the most recent experimental measurements, this mass of carbon-12 contains 6.022142×10^{23} atoms, but for most purposes, 6.022×10^{23} provides an adequate number of significant figures.



(A) The molecular mass of nitric acid is:

(a) 61 u (b) 62 u

(c) 63 u (d) 64 u

(B) Calculate the number of atoms present in 7.1 g of chlorine.

(a) 0.2 N_A (b) 0.3 N_A

(c) 0.4 N_A (d) 0.5 N_A

(C) What is the context for introducing the mole?

(D) What is the mass of 1 mole of nitrogen atoms?

(E) Assertion (A): One mole of water is equal to 6.023×10^{23} molecules.

Reason (R) : The mass of one mole of a substance in grams is called the molar mass.

(a) Both (A) and (R) are true and (R) is the correct explanation of (A).

(b) Both (A) and (R) are true but (R) is not the correct explanation of (A).

(c) (A) is true but (R) is false.

(d) (A) is false but (R) is true.

[Delhi Gov. QB 2022]

Ans. (A) *(c) 63 u*

Explanation: Molecular mass of HNO_3

= (1 × atomic mass of H) + (1 × atomic mass of N) + (3 × atomic of O).

= (1 × 1) + (1 × 14) + (3 × 16) u

= 63 u

(B) *(a) 0.2 N_A*

Explanation: We know that

1 mole of chlorine contains N_A atoms

Number of moles = (7.1/35)

= 0.2

Hence,

The number of atoms present in 7.1 g of chlorine is 0.2 N_A atoms

(C) One gram of a pure element is known to contain a large number of atoms while dealing with particles at atomic level. Thus, mole concept is a convenient method of expressing the amount of a substance.

(D) Mass of 1 mole of a substance = molar mass of that substance

Molar mass of nitrogen atom = 14 g

So, mass of 1 mole of nitrogen atoms = molar mass of nitrogen = 14 g

(E) *(b) Both (A) and (R) are true but (R) is not the correct explanation of (A).*

Explanation: Reason is not the correct explanation of above assertion since it is an independent statement.

Since,

number of water molecules

= number of moles x Avogadro's number

= 1 × 6.023 × 10^{23} molecules

TOPIC 2

PERCENTAGE COMPOSITION

Apart from knowing (dealing with) the number of entities present in a given sample, we also require an idea about the percentage of the elements present in a compound.

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

Mass % of an element in a compound

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

Consider the example of a water H_2O molecule:

The molar mass of water = 18.02 g

The atomic mass of oxygen = 16 g

The atomic mass of hydrogen = 1.01 g

$$\text{Mass \% of hydrogen} = \frac{2 \times 1.01}{18.02} \times 100$$

$$= 11.21\%$$

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

$$= 88.79\%$$

Through the percentage composition data of each element, we can calculate the molecular formula of the compound.

Empirical formula

It is the simplest formula of a compound that gives the simplest whole-number ratio of various atoms present in a compound.

For example, the empirical formula of benzene (C_6H_6) is CH and empirical formula of Hydrogen peroxide (H_2O_2) is HO.

Molecular formula

It is the formula of a compound giving the actual number of different types of atoms present in a molecule of the compound.

For example, the molecular formula of benzene is (C_6H_6).

Relation between empirical and molecular formula

Molecular formula = *n* × empirical formula

Where *n* is the simple whole number with values 1, 2, 3, 4, ___ etc.

Determination of the Empirical Formula and the Molecular Formula

The various steps involved are:

(1) Conversion of mass percent of various elements into grams.

(2) Convert mass obtained into the number of moles of each element using the formula:

$$\text{No. of moles} = \frac{\text{Mass of atom}}{\text{Molar mass of element}}$$

(3) Divide the mole value obtained by the smallest number.

If the ratio is not a whole number, then they are converted into a whole number by multiplying with a suitable integer.

(4) Write the empirical formula of the compound by mentioning the symbols of elements along with the simplest whole-number ratio of each element.

Write molecular formula using the steps mentioned below:

(1) Determine empirical formula mass by adding the atomic masses of various atoms present in the empirical formula.

(2) Divide molar mass by empirical formula mass to get integral value (n).

$$\frac{\text{Molar Mass}}{\text{Empirical formula}} = n$$

(3) Multiply empirical formula by (n) to obtain the molecular formula.

Example 2.4: A compound contains 4.07% hydrogen 24.27% carbon and 71.65% chlorine. Its molar mass is 98.96 g. What is its empirical and molecular formula? [NCERT]

Ans. Step-1

Consider the weight of the compound to be 100 g.

So, the Mass of hydrogen = 4.07g

Mass of carbon = 24.27g

Mass of chlorine = 71.65 g

Step-2

Number of moles of Hydrogen

$$= \frac{4.07\text{g}}{1.008\text{g}}$$

$$= 4.04$$

Number of moles of Carbon

$$= \frac{24.27}{12.01}$$

$$= 2.0225$$

Number of moles of Chlorine

$$= \frac{71.65\text{g}}{35.55\text{g}}$$

$$= 2.018$$

Step-3

Since 2.018 is the smallest value, divide it to get the ratio:

H	:	C	:	Cl	
$\frac{4.04}{2.018}$	:	$\frac{2.0225}{2.018}$	:	$\frac{2.018}{2.018}$	
2	:	1	:	1	Ratio

Step-4

The empirical formula of a compound = CH_2Cl

Step-5

For CH_2Cl, the empirical formula mass is:

$12.0 + (2 \times 1.008) + 35.5 = 49.5$ g

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

$$= \frac{98.96\text{g}}{49.48\text{g}}$$

$$= 2$$

We know, molecular formula = n × empirical formula

For n = 2; the molecular formula becomes $C_2H_4Cl_2$

Example 2.5: An inorganic salt gave following percentage composition on analysis:

Na = 29.11, S = 40.51, O = 30.38

Calculate the empirical formula of the salt.

Ans.

Element	Symbol	Percentage of element	At. mass of elements	Moles of elements $= \frac{\text{Percentage}}{\text{At. mass}}$	Simplest molar ratio	Whole no. molar ratio
Sodium	Na	29.11	23	$\frac{29.11}{23} = 1.266$	$\frac{1.266}{1.266} = 1$	2
Sulphur	S	40.51	32	$\frac{40.51}{32} = 1.266$	$\frac{1.266}{1.266} = 1$	2
Oxygen	O	30.38	16	$\frac{30.38}{16} = 1.899$	$\frac{1.899}{1.266} = 1.5$	3

Thus, the empirical formula is $Na_2S_2O_3$

TOPIC 3

STOICHIOMETRY AND STOICHIOMETRIC CALCULATIONS

The word 'stoichiometry' is derived from the Greek word meaning to measure an element. Stoichiometry in chemistry involves the calculation of masses (or volumes) of the reactants and products present in the chemical reaction. The relationship between the amounts of reactants and product is called the stoichiometry. The number of molecules used to balance the equation is called stoichiometric coefficient. The calculations in terms of molar masses and volumes between the different reactants and products involved in balanced chemical equations are called Stoichiometric calculations.

In the balanced chemical equation:

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

1 mol 2 mol 1 mol 2 mol

The coefficients 2 for O_2 and H_2O are called stoichiometric coefficients. The coefficient of CH_4 and CO_2 is 1 respectively.

The above chemical reaction states that

One mole of $CH_{4(g)}$ reacts with 2 moles of $O_{2(g)}$ to give 1 mole of $CO_{2(g)}$ and 2 moles of $H_2O_{(g)}$.

One mole of $CH_{4(g)}$ reacts with 2 moles of $O_{2(g)}$ to give 1 mole of $CO_{2(g)}$ and 2 moles of $H_2O_{(g)}$.

22.4 L of $CH_{4(g)}$ reacts with 44.8L of $O_{2(g)}$ to give 22.4L of $CO_{2(g)}$ and 44.8 L of $H_2O_{(g)}$.

16 grams of $CH_{4(g)}$ reacts with $2 \times 32_{(g)}$ of $O_{2(g)}$ to give 44 g of $CO_{2(g)}$ and 2×18 g of H_2O.

A general procedure is followed to carry out such calculations:

(1) Write the balanced chemical equation

(2) From the equation note the relationship between the given species and required species in terms of moles.

(3) Convert the molar values into the required parameters. For example, atomic mass, volume etc.

(4) Calculate the quantity of the desired substance by simple mathematical calculations

Limiting Reagent

The reactant that gets consumed first since it is present in a lesser amount. It limits the involvement of the other reactants in the reaction and also on the amount of product formed. therefore, it is called the limiting reagent.

Example 2.6: **3.0 g of H_2 reacts with 29.0 g of O_2 to form H_2O.**

(A) Which is the limiting reactant?

(B) Calculate the amount of H_2O formed.

(C) Calculate the amount of reactant left unreacted.

Ans. (A) $2H_2 + O_2 \rightarrow 2H_2O$

2g 32g 36g

3g of H_2 require $O_2 = \left(\frac{32}{4}\right) \times 3 = 24$ g. Thus,

O_2 (24g) is in excess. Hence, H_2 is the limiting reagent.

(B) 2 mole of hydrogen gas reacts to form 2 mole of water molecule therefore

4 g of H_2 produces = 36 g of water

So the amount of H_2O produced by 3 g

$H_2 = \left(\frac{36}{4}\right) \times 3 = 27$ g

Hence, 27 g of water will be produced during the reaction.

(C) O_2 left unreacted = 29 – 24 = 5 g

Example 2.7: **In a reaction**

$$A + B_2 \rightarrow AB_2$$

Identify the limiting reagent, if any, in the following reaction mixtures.

(A) 300 atoms of A + 200 molecules of B

(B) 2 mol A + 3 mol B

(C) 100 atoms of A + 100 molecules of B

(D) 5 mol of A + 2.5 mol of B

(E) 2.5 mol of A + 5 of mol B **[NCERT]**

Ans. (A) According to balanced equation, 1 atom of A reacts with 2 atoms of B. Thus for 300 atoms of A, 600 atoms of B are needed. Hence, B is the limiting reagent.

(B) According to balanced equation, 1 mole of A reacts with one mole of B_2 hence for 2 mol of A moles 2 moles of B_2 are needed so A is the limiting reagent.

(C) 100 atoms of A reacts with 100 molecules (200 atoms) of B, hence they are present in stoichiometric ratio.

(D) For 5 mol of A , 5 moles of B_2 are needed hence , B_2 is the limiting reagent.

(E) For 2.5 mol of A , 2.5 moles of B_2 are needed. Hence, A is the limiting reagent.

TOPIC 4

CHEMICAL REACTIONS

A chemical reaction occurs when one or more substances react to produce new substances with completely different characteristics.

Chemical Equations

A chemical equation is a simplified depiction of a chemical reaction using symbols and equations for the chemicals involved. For example, the interaction of silver nitrate with sodium chloride to produce silver chloride and sodium nitrate can be written as

$$AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$$

Reactants are the chemicals that react among themselves to cause chemical changes, whereas products are the compounds that are produced as a result of the chemical change. An arrow pointing towards the products separates the reactants and products of a chemical equation.

Characteristics of a chemical equation

The following conditions must be met by a chemical equation:

(1) It must be in accordance with the results of the experiment.

(2) It needs to include molecular species.

(3) It must be balanced, which means it must adhere to the rule of conservation of mass and have an equal amount of atoms of each element on both sides.

(1) Qualitative Information:

It describes the various reactants and products that are involved in a reaction. It gives an estimate of the number of molecules involved in or generated during the reaction.

The following adjustments can be made to the chemical equation to make it more informative:

(i) The abbreviations *(s)* for solid, *(l)* for liquid, *(g)* for gas, and *(aq)* for an aqueous solution can be used to describe the physical states of reactants and products. e.g.

$$Cu_{(s)} + 2HCl_{(aq)} \rightarrow CuCl_{2(aq)} + H_{2(g)}$$

(ii) Before the acid or base formula, dil. for dilute or conc. for concentrated is written to indicate the strength of the acid or base.

$$Cu_{(s)} + 2\ \text{dil.}HCl_{(aq)} \rightarrow CuCl_{2(aq)} + H_{2(g)}$$

(iii) Above the arrow between the reactants and products, write the reaction conditions, such as the presence of a catalyst, temperature, pressure, and so on.

$$N_{2(g)} + 3H_{2(g)} \xrightarrow[\text{800 atm}]{\text{Fe/Mo, 723 K}} 2NH_{3(g)}$$

The heat change that occurs during the reaction can be described in one of two ways.

$$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)} + 93.6\ kJ;\ \Delta H = -93.6 kJ$$

(2) Quantitative Information:

Quantitative information conveyed by a chemical equation is as follows.

(i) The proportion of reactant and product species (atoms or molecules) involved in the process.

(ii) The relative mole fractions of the reactants and products.

(iii) The mass ratios of the reactants and products.

(iv) The relative volumes of gaseous reactants and products.

Important

➥ *Old STP conditions, 273.15 K, 1 atm, volume occupied by 1 mole of a gas = 22.4 L*

New STP conditions, 273.15K, 1 bar, volume occupied by 1 mole of a gas = 22.7 L

Balancing a Chemical Equation

A chemical equation must follow the law of conservation of mass, which states that the total mass of the reactants must be equal to the total mass of the products. In a balanced chemical equation, there should be an equal amount of atoms of each element on the reactant and product sides.

For example, $Fe_{(s)} + 3O_{2(g)} \rightarrow 2Fe_2O_3$

Above reaction is a balanced chemical equation.

On the other hand, $N_2 + H_2 \rightarrow NH_3$ is an unbalanced chemical equation.

A chemical equation can be balanced by using following methods:

(1) Hit and trial method

(2) Partial equation method

(3) Ion electron method

(4) Oxidation number method

Hit and trial method

It is used to balance a chemical equation in the simplest way. It involves the following steps:

Step 1: The skeletal equation must be written using the symbols and formula of reactants and products.

Step 2: If elementary gases like hydrogen and nitrogen are present then convert them into their atomic states.

Step 3: Choose the molecular formula containing the maximum number of atoms. Then, balance the number of atoms of the species on both sides of the equation.

Otherwise, it can also be proceeded by choosing the species with the minimum number of atoms.

Step 4: After balancing the number of atoms, convert the equation into molecular form.

Step 5: Make sure the number of atoms are balanced in the final chemical equation.

Important

- *While balancing a chemical equation, the subscripts in the formula of reactants and products can't be changed.*

Partial equation method

This method is helpful only for simple reactions where a particular atom is not repeated frequently on a number of compounds. It involves the following steps:

Step 1: The given chemical reaction is broken down into various probable steps which are individually called as partial equations.

Step 2: Using the hit and trial method, each partial equation is balanced.

Step 3: Cancel out the intermediate species (if any) by multiplying partial equations with suitable integers.

Step 4: Add the partial equations to get the final chemical equation.

Reactions in Solutions

Most of the chemical reactions in the laboratory are carried out in solutions. The concentration of a solution can be expressed in several ways:

Mass per cent or weight per cent (w/w %)

It is the mass of a compound per 100 g of the solution.

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

Example 2.8: A solution is prepared by adding 2g of substance A to 18 g of water. Calculate the mass percent of the solute. [NCERT]

Ans. The mass percent of A

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

$$= \frac{2g}{2g + 18g} \times 100$$

$$= \frac{2g}{20g} \times 100$$

$= 10\%$

Mole fraction

It is the ratio of the number of moles of one component to the total number of moles of all components present in the solution. If substance A dissolves in substance B, then their mole fraction X is given as:

Mole Fraction of A

$$= \frac{\text{No. of moles of A}}{\text{No. of moles of all components in solution}}$$

$$X_A = \frac{n_A}{n_A + n_B}$$

Mole Fraction of B

$$= \frac{\text{No. of moles of B}}{\text{No. of moles of all components in solution}}$$

$$X_B = \frac{n_B}{n_A + n_B}$$

Also, the sum of mole fraction $(X_A + X_B) = 1$

Molarity (M)

It is defined as the number of moles of solute present in one litre of solution.

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

Its unit is mole per litre.

When two solutions are mixed with the same no. of moles, the general formula used is:

$$M_1V_1 = M_2V_2$$

where,

M_1 = Initial Molarity

M_2 = Molarity of new solution

V_1 = Initial Volume of solution

V_2 = The volume of new solution

Molality (m)

It is defined as the number of moles of solute present in 1 kg of solvent.

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

Unit of molality is mole per kg.

Example 2.9: The density of the 3M solution of NaCl is 1.25 g ml^{-1}. Calculate the molality of the solution.

Ans. Molarity (M) = 3 mol L^{-1}

Molar mass of NaCl = 58.5 g

Mass of NaCl in 1L solution

= Molarity × Molar mass of NaCl

= 3 × 58.5 g

= 175.5 g

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

Weight = Density × Volume
= 1.25 × 1000
= 1250g

∴ Weight of solvent = Weight of solution – the weight of solute
= 1250g – 175.5g
= 1074.5g
= 1.074 kg

So,

$$\text{Molality} = \frac{\text{No. of moles of solvent}}{\text{Mass of solvent in kg}}$$

$$= \frac{3\text{mol}}{1.0745\text{kg}}$$

= 2.79m

Caution

➥ *Students should note that the molality of a solution does not change with temperature. Since mass remains unaffected with temperature.*

Related Theory

➥ *A solution of the required concentration is made by diluting a solution of a known higher concentration. The solution of higher concentration is called a Stock solution.*

Example 2.10: Case Based:

In the health sciences, the concentration of a solution is often expressed as parts per thousand (ppt), indicated as a proportion. The labels on bottles of commercial reagents often describe the contents in terms of mass percentage. Sulfuric acid, for example, is sold as a 95% aqueous solution. Parts per million and parts per billion are used to describe concentrations of highly dilute solutions. There are several different ways to quantitatively describe the concentration of a solution. For example, molarity and molality are useful ways to describe solution concentrations for reactions that are carried out in solution. For dilute aqueous solutions, the molality and molarity are nearly the same. Mole fractions are used not only to describe gas concentrations but also to determine the vapour pressures of mixtures of similar liquids.

(A) The molarity of a solution obtained by mixing 750 mL of 0.5 M HCl with 250 mL of 2 M HCl will be:

(a) 0.975 M (b) 0.875 M
(c) 1.00 M (d) 1.175 M

[Delhi Gov. QB 2022]

(B) A molal solution is one that contains one mole of solute in:

(a) 22.4 litres of solution
(b) 1000 g of solvent
(c) 1 L of solvent
(d) 1 L of solution

(C) What is meant by mass per cent?

(D) A solution is prepared by mixing 100 mL of toluene with 300 mL of benzene. The densities of toluene and benzene are 0.86 g/mL and 0.87 g/mL, respectively. Find the mass per cent of toluene.

(E) Assertion (A): One molar aqueous solution has always a higher concentration than One molal aqueous solution.

Reason (R): The molality of the solution depends upon the density of solution whereas molarity does not.

(a) Both (A) and (R) are true and (R) is the correct explanation of (A).
(b) Both (A) and (R) are true but (R) is not the correct explanation of (A).
(c) (A) is true but (R) is false.
(d) (A) is false but (R) is true.

[Delhi Gov. QB 2022]

Ans. (A) *(b) 0.875 M*

Explanation: For mixing:

$$M_1V_1 + M_2V_2 = M_3V_3$$

$$750 \times 0.5 + 250 \times 2 = M_3\ (750+250)$$

$$M_3 = 375 + 500/1000 = 0.875\ \text{M}$$

(B) *(b) 1000 g of solvent*

Explanation:

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

Thus, one molal is one mole in 1000 g of solvent.

(C) Mass per cent: It is the percentage of the mass of solute or solvent w.r.t total mass of solution.

(D) Mass per cent of toluene

$$= \frac{\text{Mass of toluene}}{\text{Mass of toluene} + \text{Mass of benzene}} \times 100$$

Mass of toluene = Density × Volume
= 0.86 g/mL × 100 mL = 86 g

Mass of Benzene = Density × Volume

= 0.87g/mL × 300 mL

= 261 g

Mass percent of toluene

$= \frac{86}{86+261} \times 100$

$= \frac{86}{347} \times 100$

= 24.8%

(E) *(b) Both (A) and (R) are true but (R) is not the correct explanation of (A).*

Explanation: One molar aqueous solution is more concentrated than one molal aqueous solution because 1 M solution means 1 mole of solute dissolved in 1 litre of the solution which include both solute and solvent. So, the mass of solvent (i.e. water) is less than 1000 gram. Therefore 1 m aqueous solution contains 1 mole of solute dissolved in less than 1000 gram of solvent whereas 1 molal solution has 1 mole of solute in 1000 gram of solvent. Hence concentration will be more in 1 molar aqueous solution.

OBJECTIVE Type Questions

[1 mark]

Multiple Choice Questions

1. One mole of a substance has 6.022×10^{23} atoms/molecules. Find the number of molecules present in 100 mL of 0.02 M H_2SO_4 solution.
   (a) 12.044×10^{20} molecules
   (b) 6.022×1023 molecules
   (c) 1×10^{23} molecules
   (d) 12.044×10^{23} molecules

[NCERT Exemplar]

Ans. *(a) 12.044×10^{20} molecules*

Explanation:

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

$$0.02 = \frac{\text{No. of moles}}{\frac{100}{1000\text{ L}}}$$

Number of moles

$= 0.02 \times 0.1$

$= 0.002$ mol

No. of molecules of H_2SO_4

= No. of moles of $H_2SO_4 \times N_A$

$= 0.002 \times 6.022 \times 10^{23}$

$= 12.044 \times 10^{20}$ molecules

2. Find the correct match:

Column X	Column Y	Column Z
(A) 40 g of He	(i) 3.011×10^{23} atoms	(p) 0.5 moles
(B) 35 g of Li	(ii) 10 atoms	(q) 1.67×10^{-23} moles
(C) 40 u of He	(iii) 6.022×10^{23} atoms	(r) 10 moles
(D) 16 g of O_2	(iv) 3.011×10^{24} atoms	(s) 5 moles

(a) A (ii) (p)
(b) B (iv) (s)
(c) C (iii) (q)
(d) D (i) (r)

[Delhi Gov. QB 2022]

Ans. *(b) B (iv) (s)*

Explanation: 40 g of He → Mole = 40/4 = 10 → $10 \times 6.022 \times 10^{23}$ atoms = 6.022×10^{24} atoms (A-iii-r)

35 g of Li → Mole = 35/7 = 5 mole → $5 \times 6.022 \times 10^{23}$ atoms = 3.011×10^{24} atoms (B-iv-s)

40 u of He → 40/4=10 atoms → Moles = $10/6.022 \times 10^{23}$ = 1.67×10^{-23} moles (C-ii-q)

16 g of O_2 → 16/32 = 0.5 mole → $0.5 \times 6.022 \times 10^{23}$ atoms = 3.011×10^{23} atoms (D-i-p)

3. The empirical formula and molecular mass of a compound are CH_2O and 180 g respectively. Find the molecular mass of the compound.
   (a) CH_2O
   (b) $C_9H_{18}O$
   (c) $C_6H_{12}O_6$
   (d) $C_2H_2O_2$

[NCERT Exemplar]

Ans. *(c) $C_6H_{12}O_6$*

Explanation: Empirical formula mass of CH_2O

$= 12 + (2 \times 1) + 16 = 30$ g

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

$$= \frac{180\text{g}}{30\text{g}}$$

= 6

Now, Molecular formula = $6 \times (CH_2O)$

$= C_6H_{12}O_6$

Related Theory

- *If n = 1 then the empirical and molecular formulas are the same, otherwise different.*

  The compounds are known only by their molecular formula and not by the empirical formula.

4. One mole of CH_4 contains:

(a) 4 g atoms of hydrogen

(b) 3 g atoms of carbon

(c) 6.022×10^{23} atoms of hydrogen

(d) 1.83×10^{23} molecules of CH_4

Ans. *(a) 4 g atoms of hydrogen*

Explanation: 1 mole of methane (CH_4) has = 4 moles of H atoms = 4 g atoms of H atoms

1 mole of methane (CH_4) has = 1 mole of C atom = 12 g-atoms of C atom

1 mole 1 mole of methane (CH_4) has

$= 4 \times 6.02 \times 10^{23}$ of H atoms

$= 24.08 \times 10^{23}$ of H atoms.

5. Calculate the atomic weight of an element A present in a sample containing 2.64×10^{22} atoms weighing 2.15g:

(a) 35 g **(b) 49 g**

(c) 55 g **(d) 42 g**

Ans. *(b) 49 g*

Explanation: Number of atoms in a sample = 2.64×10^{22}

Weight of sample = 2.15 g

Number of atoms = number of moles × atomicity × N_A

$$= 2.64 \times 10^{22}$$

$$= \frac{\text{Given weight}}{\text{Atomic weight}} \times 1 \times 6.022 \times 10^{23}$$

$$\text{Atomic weight} = \frac{2.15 \times 6.022 \times 10^{23}}{2.64 \times 10^{22}} = 49 \text{ g}$$

Related Theory

- *Atomicity is defined as the total number of atoms present in a molecule.*

  For example, the atomicity of oxygen (O_2) is 2

6. What is the percentage composition by mass of carbon present in urea?

(a) 25% **(b) 28%**

(c) 42% **(d) 20%**

Ans. *(d) 20%*

Explanation: The molecular formula of Urea $\rightarrow \left[CO(NH_2)_2\right]$.

The total mass of the compound

$= (12 + 16 + 2 \times 14 + 4 \times 1)$g

= 60 g

Mass of carbon atom = 12 g

Therefore, the mass percentage of carbon

$$= \frac{12 \text{ g}}{60 \text{ g}} \times 100$$

= 20%

7. From 120 g of zinc sulphate ($ZnSO_4$), the amount of zinc that can be obtained is:

(a) 45.0 g Zn **(b) 41.0 g Zn**

(c) 48.58 g Zn **(d) 39.9 g Zn**

Ans. *(c) 48.58 g Zn*

Explanation: Molar mass of the compound ($ZnSO_4$)

$= 65.3 + 32 + 16 \times 4$

= 161.3 g

So, 161.3 g ($ZnSO_4$) contain = 65.3 g Zn

Then, 120 g ($ZnSO_4$) contain

$$= \frac{65.3}{161.3} \times 120$$

= 48.58 g of Zn

8. 10 g of $MgCO_3$ is dissolved in enough water to obtain 300 mL of the solution, what will be the molarity of $MgCO_3$ in the solution?

(a) 0.39 M **(b) 0.52 M**

(c) 0.76 M **(d) 0.1 M**

Ans. *(a) 0.39 M*

Explanation: Molar mass of $MgCO_3$

$= (24.3 + 12 + 16 \times 3)$g

= 84.3 g

Given mass of $MgCO_3$ = 10g

Volume of solution

= 300 mL

$$= \frac{300}{1000}$$

= 0.3 litres

$$\text{Molarity} = \frac{\frac{\text{Given mass}}{\text{molar mass}}}{\text{Volume of solution in litre}}$$

$$= \frac{\frac{10}{84.3}}{0.3}$$

= 0.39 M

Assertion-Reason (A-R)

In the following question no. (9-12), a statement of assertion followed by a statement of reason is given. Choose the correct answer out of the following choices:

(a) Both (A) and (R) are true and (R) is the correct explanation of (A).

(b) Both (A) and (R) are true but (R) is not the correct explanation of (A).

(c) (A) is true but (R) is false.

(d) (A) is false but (R) is true.

9. Assertion (A): $NaHCO_3$ has the same empirical and molecular formula.

Reason (R): $NaHCO_3$ evolves CO_2 gas on heating.

Ans. *(b) Both (A) and (R) are true but (R) is not the correct explanation of (A).*

Explanation: Molecular formula

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

The value of n

$$= \frac{\text{Molar mass of } NaHCO_3}{\text{Empirical formula mass } NaHCO_3}$$

$$= \frac{84\ g/mol}{84\ g/mol}$$

$$= 1$$

So, the empirical formula is equal to the molecular formula and this has no relation with the fact that $NaHCO_3$ evolves CO_2 gas on heating.

10. Assertion (A): 8 g of CH_4 and 32 g of SO_2 have the same number of molecules.

Reason (R): Two substances with an equal number of moles have an equal number of molecules.

Ans. *(a) Both (A) and (R) are true and (R) is the correct explanation of (A)*

Explanation: Number of moles of CH_4

$$= \frac{\text{Mass}}{\text{Molar mass}}$$

$$= \frac{8\ g}{16\ g} = 0.5$$

Number of moles of SO_2

$$= \frac{\text{Mass}}{\text{Molar mass}}$$

$$= \frac{32\ g}{64\ g} = 0.5$$

As both compounds have the same number of moles, so they have the same number of molecules.

11. Assertion (A): Combustion of 16 g of methane gives 18 g of water.

Reason (R): In the combustion of methane, water is one of the products.

[NCERT Exemplar]

Ans. *(d) (A) is false but (R) is true.*

Explanation: $CH_4 + 2O_2 \rightarrow CO_{2(g)} + 2H_2O_{(g)}$

16g 36g

So, the combustion of 16 g of CH_4 gives 36g of water.

12. Assertion (A): Molarity and molality of a solution do not depend on temperature.

Reason (R): Molarity and molality depend on the number of moles of solute particles.

Ans. *(d) (A) is false but (R) is true.*

Explanation: The molarity of a solution depends upon the temperature whereas molality does not. This is due to the reason that the molarity of a solution is related to the volume of solution which in turn is related to temperature, whereas molality is related to the mass of solvent.

CASE BASED Questions (CBQs)

[4 & 5 marks]

Read the following passages and answer the questions that follow:

13. *A binary solution is made up of two liquids that are entirely miscible with each other. In a binary solution, the component with the lowest concentration is known as the solute, while the component with the highest concentration is known as the solvent. One mole of the solute present in one litre of solution is referred to as a 1 molar solution. A 1 molal solution is one in which one mole of solute is dissolved in one kilogram of solvent. The number of moles of a given component to the total number of moles in the solution is referred to as the mole fraction.*

(A) 6.02×10^{20} molecules of urea are present in 100 mL of its solution. The concentration of the solution is:

(a) 0.02 M (b) 0.01 M
(c) 0.001 M (d) 0.1 M

[Delhi Gov. QB 2022]

(B) What will be the mole fraction of glycol $C_2H_4(OH)_2$ in a solution containing 45 g of water and 56 g of glycol?
(a) 0.31 (b) 0.50
(c) 0.26 (d) 0.10

(C) The value of molality for pure water is:
(a) 55.55 (b) 52.6
(c) 52 (d) 25

(D) What is the mass per cent of the carbon in ethanol?
(a) 59 (b) 42
(c) 45 (d) 52

[Delhi Gov. QB 2022]

(E) What is the correct advantage for using molality over molarity?
(a) Molarity does not depend upon temperature.
(b) Molality does not depend upon temperature.
(c) Molality depend on temperature.
(d) None of the above

Ans. (A) *(b) 0.01 M*

Explanation:

$$\text{Number of moles} = \frac{\text{Molecules of urea}}{\text{Avogadro's number}}$$

$$= \frac{6.02\times10^{20}}{6.02\times10^{23}}$$

$$= 10^{-3}$$

Molarity = Number of moles of solute/volume of solution

$$= \frac{10^{-3}}{0.1} = 0.01\text{M}$$

(B) *(c) 0.26*

Explanation: Mole fraction of glycol

$$= \frac{\text{No. of moles of glycol}}{\text{No. of moles glycol} + \text{No. of moles of water}}$$

$$= \frac{\frac{56}{62}}{\frac{56}{62}+\frac{45}{18}}$$

$$= \frac{0.9}{0.9+2.5} = 0.26$$

(C) *(a) 55.55*

Explanation: Molality

$$= \frac{\text{No. of moles of solute}}{\text{Mass of solvent in kg}}$$

Molality for a water molecule

$$= \frac{\frac{45}{18}}{\frac{45}{1000}} = 55.55 \text{ m}$$

(D) *(d) 52*

Explanation: Molecular mass of ethanol

$= 2 \times 12 + 6 \times 1 + 16$

$= 46\ u$

The mass per cent of carbon

$$= \frac{\text{Mass of Carbon}}{\text{Mass of Ethanol}} = \frac{24}{46} \times 100 = 52\%$$

(E) *(b)* Molality does not depend upon temperature.

Explanation: Molality is favored over molarity as the unit of concentration because molality is a function of temperature and changes with temperature but molarity is independent of temperature so it stays the same. The mass of the solvent is also independent of temperature so it remains constant.

14. *The reactants react according to the balanced chemical equation. Quite often, these are not present in the same proportions as is required by the equation; some may be present in a lesser amount while the others may be present in excess than the stoichiometric amounts. The reactant which is present in a lesser quantity is known as a limiting reagent or limiting reactant since it limits the participation of the other reactants in the reaction and also the product of the reaction. For example, in the combustion of methane with oxygen methane is the limiting reactant because oxygen is always available more than the amount of methane. Amount of carbon dioxide and water formed also depends upon the amount of methane and not oxygen.*

(A) Find the number of moles of lead (II) chloride formed as a result of the reaction between 6.5 g of PbO and 3.2 g of HCL.

(B) 14g hydrogen and 80 g oxygen were filled in a steel vessel and exploded. The amount of water produced in the reaction will be?

(C) Why is it necessary to balance a chemical equation?

Ans. (A)

PbO	+	$2HCl$	→	$PbCl_2$	+	$2H_2O$
1 mole		2 moles		1 mole		2 moles
$\frac{6.5}{224}$ mol		$\frac{3.2}{36.5}$ mol				
0.029 mol		0.087 mol				

So, PbO is the limiting reactant.

= 0.029 mol of $PbCl_2$ is formed.

(B) $H_{2(g)} + \frac{1}{2}O_{2(g)} \rightarrow H_2O_{(l)}$

1 mole 0.5 mole 1 mole

14 g of $H_2 = \frac{14}{2}$ mole $= 7$ mol

80 g of O_2

$= \frac{80}{32}$ mole $= 2.5$ mol

So, O_2 is the limiting reagent.

Since 0.5 mole of oxygen from water = 1 mol

So, 2.5 mol of oxygen form water

$= \frac{1}{0.5} \times 2.5$

$= 5$ mol

(C) A chemical equation has to be balanced in order to satisfy the law of conservation of mass. According to the law, there is no change in mass when the reactants change into the products. Therefore, the chemical equation has to be balanced.

VERY SHORT ANSWER Type Questions (VSA)

[1 mark]

15. What is the SI unit of molality?

Ans. SI unit of molality is mol per kg.

Related Theory

- *Molality does not depend on temperature. This is because molality of the solution is dependent on the mass of the solvent and mass is independent of temperature.*

16. What do you mean by stoichiometric coefficients in a chemical equation?

Ans. In a balanced chemical equation, the balancing coefficients of reactants and products are known as stoichiometric coefficients.

For example,

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

The stoichiometric coefficients are 1, 3 and 2 respectively.

Related Theory

- *The stoichiometric calculations involve three cases:*
  (1) Mass-Mass relationship: In this case, the mass of one of the reactants or products is given and that of the other can be calculated.
  (2) Mass-Volume relationship: In this situation, one of the reactants' or products' mass or volume is known, and the other can be calculated.
  (3) Volume-Volume relationship: In this case, the volume of one of the reactants or products is given and that of the other can be calculated.

17. What is the symbol for the SI Unit of a mole? How is the mole defined? [NCERT Exemplar]

Ans. The symbol for the SI Unit of a mole is mol.

A mole is defined as a collection of Avogadro's number or 6.022×10^{23} particles (atoms, ions and molecules).

Related Theory

- *1 mole of carbon atoms $= 6.022 \times 10^{23}$ carbon atoms*
- *1 mole of oxygen molecules $= 6.022 \times 10^{23}$ oxygen molecules*
- *1 mole of water molecules $= 6.022 \times 10^{23}$ water molecules*
- *1 mole of electrons $= 6.022 \times 10^{23}$ electrons.*

Caution

- *While using the term mole, students should always specify the kind of particles involved.*

SHORT ANSWER Type-I Questions (SA-I)

[2 marks]

18. 5.8 g butane burns in the presence of oxygen to produce carbon dioxide and water vapour. Calculate the mass of carbon dioxide produced. [Diksha]

Ans. $2C_4H_{10(g)} + 13O_{2(g)} \rightarrow 8CO_{2(g)} + 10H_2O_{(l)}$

2 moles 13 moles 8 moles 10 moles

116 g 352 g

Given mass of butane = 5.8 g

Since, 116 g of butane produces

= 352 g of CO_2

So, 5.8g of butane will produce

$= \frac{352}{116} \times 5.8$

$= 17.6$ g of CO_2

19. How much carbon disulphide will be formed if 4 g of carbon was heated with 8 g of sulphur. Also, identify the limiting reagent.

Ans. $C \quad + \quad 2S \rightarrow CS_2$

12 g   64 g   76 g

Sulphur will be the limiting reactant.

8 g sulphur will produce $CS_2 = \frac{76}{64} \times 8\ g = 9.5\ g$

20. Water has a density of 1.0g/mL at room temperature. If the volume of a drop of water is 0.05 mL. How many molecules are there?

Ans. The volume of a drop of water = 0.05 mL

Mass of a drop of water

$= \text{Volume} \times \text{density}$

$= (0.05\ mL) \times (1.0 g/mL)$

$= 0.05 g$

The molecular mass of water (H_2O)

$= 2 \times 1 + 16 = 18\ g$

We know, 18 g of water = 1 mol

Hence, 0.05 g of water will be $= \frac{1\ mol}{18\ g} \times 0.05\ g$

$= 0.0028\ mol$

Now, 1 mole of water contains

$= 6.022 \times 10^{23}$ 0.0028 moles of water contain

$= 6.022 \times 10^{23} \times 0.0028$

$= 1.68 \times 10^{21}$ molecules

21. A compound has 9 carbon atoms and 13 hydrogen atoms and mass of the remaining entity is 2.33×10^{-23}g.

Find the molecular mass of the compound.

Ans. Mass of 9 carbon atoms

$= 9 \times 12\ u$

$= 108\ u$

Mass of 13 hydrogen atoms

$= 13 \times 1\ u$

$= 13\ u$

Mass of 2.33×10^{-23}g other entities

$$= \frac{(2.33 \times 10^{-23} g)}{(1.66 \times 10^{-24} g)}$$

$= 14.04\ u$

The molecular mass of the substance

$= (108 + 13 + 14.04)\ u$

$= 135.04\ u$

22. Calculate the molarity of NaOH in the solution prepared by dissolving 4 g in enough water to form 25 mL of the solution. **[NCERT Exemplar]**

Ans. $\text{Molarity(M)} = \frac{\text{No. of moles of solute}}{\text{volume of solution in litres}}$

Also, no. of moles of solute

$= \frac{\text{Mass of NaOH}}{\text{Molar mass of NaOH}}$

$= \frac{4 g}{40 g}$

$= 0.1\ mol$

So, Molarity

$= \frac{0.1\ mol}{0.025 L}$

$= 4\ M$

⚠ Caution

➥ *While solving a problem, students should always remember that during molarity calculation, volume must be converted into litre.*

23. Hydrogen gas is prepared in the laboratory by reacting dilute HCl with granulated zinc. Following reaction takes place:

$$Zn_{(s)} + 2HCl_{(aq)} \xrightarrow{aq} ZnCl_{2(aq)} + H_{2(g)}$$

Calculate the volume of hydrogen gas liberated at STP when 32.65 g of zinc reacts with HCl. Given that one mole of a gas occupies 22.7 L volume at STP; the atomic mass of Zn = 65.3 u. **[NCERT Exemplar]**

Ans. $Zn_{(s)} + 2HCl_{(aq)} \xrightarrow{aq} ZnCl_{2(aq)} + H_{2(g)}$

(65.3 u)   22.7L

65.3 g of zinc evolve hydrogen gas at STP = 22.7 L 32.65 g of zinc evolve hydrogen gas at STP

$= \frac{32.65\ g}{65.3\ g} \times 22.7\ L$

$= 11.35\ L$

Related Theory

➥ *STP refers to Standard Temperature and Pressure. It is the experimental condition at which temperature is maintained at 273 K and 1 atm pressure.*

24. In the reaction 2A + 4B → 3C + 4D, when 5 moles of A react with 6 moles of B, then

(A) Which is the limiting reagent?

(B) Calculate the amount of C formed.

[NCERT Exemplar]

Ans. (A) The given reaction is:

$2A \quad + 4B \rightarrow \quad 3C \quad + \quad 4D$

(2 mol) (4 mol)   (3 mol)   (4 mol)

2 moles of A react with moles of B = 4

5 moles of A react with 10 moles of B. But moles of B available = 6 mol.

This means that B is the limiting reagent.

(B) Amount of C formed

$$= \frac{(3\ mol)}{(4\ mol)} \times (6\ mol)$$

$= 4.5\ mol$

SHORT ANSWER Type-II Questions (SA-II)

[3 marks]

25. Find the mass percentage of each component in the following compound: calcium phosphate.

Ans. Molecular mass of $Ca_3(PO_4)_2$

= 3 × (Atomic mass of Ca) + 2 × atomic mass of P + 8 × atomic mass of O

= (3 × 40 u) + (2 × 31 u) + (8 × 16 u)

= 310 u

The mass per cent of calcium (Ca)

$$= \frac{(120\ u)}{(310\ u)} \times 100 = 38.71\%$$

The mass per cent of phosphorus (P)

$$= \frac{(62\ u)}{(310\ u)} \times 100 = 20\%$$

The mass per cent of oxygen (O)

$$= \frac{(128\ u)}{(310\ u)} \times 100 = 41.29\%$$

26. A vessel contains 1.6 g of dioxygen at STP (273 K, 1 atm pressure). The gas is now transferred to another vessel at a constant temperature where pressure becomes half of the original pressure. Calculate:

(A) volume of the new vessel

(B) number of molecules of dioxygen.

[NCERT Exemplar]

Ans. (A) 32 g of oxygen at STP occupies volume = 22.4L

1.6 g of oxygen at STP occupy volume

$$= (22.4\ L)\frac{(1.6\ g)}{(32.3\ g)}$$

= 1.12 L

Given

V_1 = 1.12 L

Let P_1 = P atm

and $P_2 = \frac{P}{2}$ atm;

Applying Boyle's Law. $P_1V_1 = P_2V_2$

$$V_2 = \frac{P_1 \times V_1}{P_2}$$

$$= \frac{(P_{atm}) \times (1.12\ L)}{\frac{P}{2}atm}$$

= 2.24 L

(B) 32 g of oxygen contain = 6.022 × 10^{23} molecules.

So. 1.6 g of oxygen contain

$$= (6.022 \times 10^{23}) \times \frac{(1.6\ g)}{(32\ g)}$$

= 3.011 × 10^{22} molecules

27. To complete a reaction, 0.184 g of NaOH must be added to a reaction vessel. For this reaction, how many millilitres of 0.150 M NaOH a solution should be added?

Ans. Mass of NaOH = 0.184 g

Molar mass of NaOH = 40.0 g mol^{-1}

Molarity of solution

= 0.150 m

= 0.15 mol L^{-1}

Let the Volume of the solution = V mL

Now, the Molarity of solution

$$M = \frac{\text{Molar mass}}{\text{Volume of solution}}$$

$$0.15\ mol\ L^{-1} = \frac{\frac{\text{Mass of NaOH}}{\text{Molar mass}}}{\text{Volume of solution}}$$

Volume of Solution

$$= \frac{(0.184\ g)}{(40 g\ mol^{-1}) \times \text{Volume of solution}}$$

$$= \frac{(0.184\ g)}{(40 g\ mol^{-1}) \times (0.15\ mol\ L^{-1})}$$

= 0.0307 L

= 30.7 mL

LONG ANSWER Type Questions (LA)

[4 & 5 marks]

28. 3.1 g of the carbohydrate on heating in the absence of oxygen yields 1.24 g of carbon. The molecular mass of the carbohydrate is 180 u. Calculate its molecular formula. Find the molecular formula of a carbohydrate moiety whose molar mass is given as 180 u. If it is heated in absence of oxygen, it produces 1.24 g of carbon. Given the weight of the sample is 3.1 g.

Ans. Carbohydrate has the general formula: $C_x(H_2O)_y$

Percentage of carbon (C) in carbohydrate

$$= \frac{(1.24\text{ g})}{(3.10\text{ g})} \times 100$$

$$= 40\%$$

Percentage of water (H_2O) in carbohydrate = 100 – 40 = 60%.

Number of moles of Carbon

$$= \frac{40\text{ g}}{12\text{ g mol}^{-1}}$$

$$= 3.33\text{ mol}$$

Number of moles of H_2O

$$= \frac{60\text{ g}}{18\text{ g mol}^{-1}}$$

$$= 3.33\text{ mol}$$

Mole ratio of C : H_2O = 1 : 1

The empirical formula of carbohydrate = $C(H_2O)$

Empirical formula mass = 12 + 18 = 30 u

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

$$= \frac{180\text{ g}}{30\text{ u}}$$

$$= 6$$

The molecular formula of carbohydrate

$= 6 \times C(H_2O)$

$= C_6H_{12}O_6$

29. 0.38 gram of NaOH was weighed and its aqueous solution was prepared. The solution was made up to 50 mL in a volumetric flask.

(A) What is the molarity of the final solution?

(B) Find the number of moles of NaOH which are present in 27 mL of 0.15 M NaOH solution.

Ans. (A) Molarity of solution (M)

$$= \frac{\text{Number of moles of NaOH}}{\text{Volume of solution in litres}}$$

$$= \frac{\frac{0.38\text{ g}}{40\text{ g mol}^{-1}}}{\frac{50}{1000\text{ L}}}$$

$$= \frac{0.38 \times 1000}{40 \times 50}(\text{mol L}^{-1})$$

$$= 0.19\text{ mol/L}$$

(B) Given Molarity of the solution (M)

$= 0.15\text{ mol L}^{-1}$

Volume of solution

$= 27\text{ mL}$

$$= \frac{27}{1000}$$

$= 0.027\text{ L}$

Molarity of solution (M)

$$= \frac{\text{No. of moles of NaOH}}{\text{Volume of solution in litres}}$$

$$(0.15\text{ mol L}^{-1}) = \frac{\text{No. of moles of NaOH}}{0.027\text{ L}}$$

So, the number of moles of NaOH

$= (0.15\text{ mol L}^{-1}) \times (0.027\text{ L})$

$= 4.05 \times 10^{-3}\text{ mol}$

30. Calcium carbonate reacts with aqueous HCl to give $CaCl_2$ and CO_2 according to the reaction given below:

$$CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$$

What mass of $CaCl_2$ will be formed when 250 mL of 0.76 M HCl reacts with 1000 g of $CaCO_3$? Name the limiting reagent. Calculate the number of moles of $CaCl_2$ formed in the reaction. [NCERT Exemplar]

Ans. Molarity of solution (M)

$$= \frac{\frac{\text{Mass of HCl}}{\text{Molar mass of HCl}}}{\text{Volume of solution in litres}}$$

$$0.76\text{ mol L}^{-1} = \frac{\frac{(\text{Mass of HCl})}{(36.5\text{ g mol}^{-1})}}{\frac{(250)}{(1000\text{ L})}}$$

Mass of HCl = 0.25 × 0.76 × 36.5 × 1000

= 6935

$$CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + CO_{2(g)} + H_2O_{(g)}$$

(100 g) (2 × 36.5 = 73 g) (111 g)

100g of $CaCO_3$ need HCl = 73g

100g of $CaCO_3$ need HCl

$$= \frac{(73g)}{(100g)} \times (1000g)$$

= 730 g

But the amount of HCl actually available = 6.935 g

∴ HCl is the limiting reagent

73 g of HCl form $CaCl_2$ = 111 g

6.935 g of HCl form $CaCl_2$

$$= \frac{(111g)}{(73g)} \times (6.935)$$

= 10.54 g

Moles of $CaCl_2$ formed

$$= \frac{\text{Mass of } CaCl_2}{\text{Molar mass of } CaCl_2}$$

$$= \frac{(10.54g)}{(111 g mol^{-1})}$$

= 0.095 mol

31. H_2SO_4 has a molality of 0.8 M and a density of 1.06 g cm^{-3}. In terms of molality and mole fraction, what will the solution's concentration be?

Ans. Given

Molarity, $M = 0.8$ M

Density, $d = 1.06$ g/cm^3

Mass of one litre of solution

$= V \times d$

$= 1000$ cm^3 × 1.06 g cm^{-3}

= 1060 g

Mass of H_2SO_4 in 0.8 M solution

$$0.8 = \frac{\frac{\text{Mass}}{98}}{1L}$$

Mass = 98 × 0.8 = 78.4 g

Mass of Solvent (H_2O)

= (1060 – 78.4) = 981.6g

The molality of solution (m)

$$= \frac{\frac{\text{Mass of } H_2SO_4}{\text{Molar mass}}}{\text{Mass of solvent in kg}}$$

$$= \frac{\frac{(78.4\ g)}{(98 g mol^{-1})}}{\frac{(981.6)}{(1000\ kg)}}$$

= 0.815 mol kg^{-1}

= 0.815 m

Mole fraction of solute

$$= \frac{\text{No. of moles of } H_2SO_4}{\text{No. of moles of } H_2SO_4 + \text{No. of moles of water}}$$

$$= \frac{0.8\ mol}{0.8 + \frac{981.6}{18} mol}$$

$$= \frac{0.8\ mol}{(0.8 + 54.53)}$$

= 0.0145

32. Answer the following questions:

(A) In a compound $C_xH_yO_z$, the mass % of C and H is 6 : 1 and the amount of oxygen present is equal to the half of the oxygen required to react completely C_xH_y. Find the empirical formula of the compound.

(B) If the density of methanol is 0.793 kg L^{-1}, what is the volume needed for making 2.5 L of its 0.25 M solution?

Ans. (A) Moles of C $= \frac{6}{12} = \frac{1}{2}$

Moles of Hydrogen $= \frac{1}{1} = 1$

Reaction of Oxygen with

$$C_xH_y + \left(x + \frac{y}{4}\right)O_2 \rightarrow xCO_2 + \frac{y}{2}H_2O$$

Moles of oxygen (O_2) needed

$$= \frac{1}{2} + \frac{1}{4} = \frac{3}{4} \text{ moles}$$

Thus oxygen (O_2) present $= \frac{3}{8}$ moles

Ratio : C : H : O :: $\frac{1}{2} : 1 : \frac{3}{4}$ = 2 : 4 : 3

Thus empirical formula is $C_2H_4O_3$

(B) 0.25 M solution means 0.25 moles of methanol are present in 1 L solution. Thus, 2.5 L will contain = 0.25 × 2.5 = 0.625 mole

Mass of methanol = Moles × Molar mass

= 0.625 × 32 = 20 g

$$\text{Volume} = \frac{\text{Mass}}{\text{Density}} = \frac{0.020}{0.793}$$

= 0.02522 L

= 25.22 mL