1. Some Basic Concepts of Chemistry

Exercises

1. Choose the most correct option. (Answer are given directly)

A. A sample of pure water, whatever the source always contains by mass of oxygen and 11.1 % by mass of hydrogen.

Ans. 88.8

B. Which of the following compounds can NOT demonstrate the law of multiple proportions ?

Ans. Na_2S, NaF

C. Which of the following temperature will read the same value on celsius and Fahrenheit scales.

Ans. -40°

D. SI unit of the quantity electric current is...

Ans. Ampere

E. In the reaction $N_2 + 3H_2 \rightarrow 2NH_3$, the ratio by volume of N_2, H_2 and NH_3, is 1 : 3 : 2 This illustrates the law of..

Ans. gaseous volumes

F. Which of the following has maximum number of molecules ?

Ans. 2g H_2

G. How many g of H_2O are present in 0.25 mol of it?

Ans. 4.5

H. The number of molecules in 22.4 cm^3 of nitrogen gas at STP is..

Ans. 6.022×10^{20}

I. Which of the following has the largest number of atoms?

Ans. 1g Li (s)

2. Answer the following questions.

A. State and explain Avogadro's law.
Ans. Avogadro's law: It states that equal volumes of different gases under identical conditions of pressure and temperature contain equal number of molecules.

Explanation: Consider the reaction between gaseous hydrogen and oxygen forming water vapour.

$2H_{2(g)}$	$+ \; O_{2(g)}$	$\longrightarrow 2H_2O_{(g)}$
100 mL	50 mL	100 mL
2 volumes	1 volume	2 volumes

Since volume V of a gas is directly proportional to number of gas molecule ($V \alpha n$),
2 n molecules
n molecule
2 n molecules
OR
2 molecules
1 molecule
2 molecules
Thus 2 molecules of hydrogen combine with 1 molecule of oxygen to give 2 molecules of water vapour.

B. Point out the difference between 12 g of carbon and 12 u of carbon

Ans. 12 g of carbon represents one mole of carbon containing 6.022 x 10^{23} atoms of carbon.
12 u of carbon represents one carbon atom.

C. How many grams does an atom of hydrogen weigh?

Ans. Atomic mass of hydrogen is 1.0 u. Hence its mass is 1.0 × 1.66054 x 10^{-24} g = 1.66054 x 10^{-24} g.

D. Calculate the molecular mass of the following in u.

a. NH_3

b. CH_3COOH
c. C_2H_5OH

Ans. Molecular mass:
(a) NH_3 = 14 u + 3 x 1 u = 17 u.

(b) CH_3COOH = 2 x 12 u + 4 x 1u + 2 x 16 u = 60 u.
(c) C_2H_5OH

→ = 2 x 12 u + 6 x 1 u + 16 u = 46 u.

E. How many particles are present in 1 mole of a substance?

Ans. One mole of a substance contains 6.022 x 10^{23} particles.

E. What is the SI unit of amount of a substance?

Ans. The SI unit of amount of a substance is 'mol'.

G. What is meant by molar volume of a gas?

Ans. The volume of one mole of a gas is called molar volume. At STP condition (0 °C and 1 atm), the molar volume of any gas is 22.4 dm^3.

H. State and explain the law of conservation of mass.

Ans. Law of conservation of mass: This law states that during chemical combination of matter, the mass is neither created nor destroyed.

Explanation: Lavoisier while studying several combustion reactions, determined accurately the masses of the reactants before reactions and the masses of the products after the completion of the reactions. He observed that the total masses of the reactants before the reaction were in the agreement with the total masses of the products.
For example,

C	+	O_2	$\longrightarrow$	CO_2
(s)		(g)		(g)
(12 g)		(32 g)		(44 g)

The total mass of carbon (12 g) and oxygen taken initially (32 g) is equal to the mass of Carbon dioxide (44 g) formed.

I. State the law of multiple proportions.

Ans. The law of multiple proportions: This law states that when two elements chemically combine to form two or more compounds with different compositions by weights, then the masses of one element that combine with a fixed mass of the other element are in the ratio of small whole numbers.

3. Give one example of each

A. homogeneous mixture

Ans. An aqueous solution of sugar is an example of a homogeneous mixture.

B. heterogeneous mixture

Ans. A mixture of oil and water is an example of heterogeneous mixture.

C. element

Ans. Cu is an Element

D. compound

Ans. Water, H_2O is a compound

4. Solve problems:

A. What is the ratio of molecules in 1 mole of NH_3 and 1 mole of HNO_3.

Ans. 1 mole of NH_3 and 1 mole of HNO_3 contain 6.022×10^{23} molecules each. Hence the ratio of molecules is 1 : 1.

B. Calculate number of moles of hydrogen in 0.448 litre of hydrogen gas at STP

Solution: At STP 22.4 litre of H_2 contains 1 mol of H_2
∴ At STP 0.448 litre of H_2 will contain,

$$\frac{0.448}{22.4} = 0.02 \text{ mol } H_2$$

Ans. Moles of $H_2 = 0.02$ mol.

C. The mass of an atom of hydrogen is 1.008 u. What is the mass of 18 atoms of hydrogen. (18.144 u)

Solution: Mass of hydrogen atoms = $18 \times 1.008 = 18.144$ u
Ans. Mass of hydrogen atoms = 18.144 u.

D. Calculate the number of atom in each of the following (Given: Atomic mass of I = 127 u).

a. 254 u of iodine (I)

b. 254 g of iodine (1)

Solution: (a) Number of atoms of iodine (I)

$= \frac{254}{127} = 2\, atoms$

(b) Mass of 1 mol of I atoms = 127 x 6.022 x 10^{23} u = 127 × 6.022 × 10^{23} × 1.66054 x 10^{-24} g = 127 g mol^{-1}

∴ Number of moles of I

$= \frac{254}{127}$ x 2 mol

∴ Number of I atoms = 2 x 6.022 × 10^{23}

= 1.2044 x 10^{24} atoms

Ans. (a) Number of I atoms = 2

(b) Number of I atoms = 1.2044 x 10^{24}

E. A student used a carbon pencil to write his homework. The mass of this was found to be 5 mg. With the help of this calculate.

a. The number of moles of carbon in his homework writing.

b. The number of carbon atoms in 12 mg of his homework writting.

Solution: (a) Number of moles of carbon

$= \frac{5 \times 10^{-3}}{12} = 4.166 \times 10^{-4} mol$

(b) Number of moles carbon

$= \frac{12 \times 10^{-3}}{12} 1 \times 10^{-3} mol$

∴ Number of carbon atoms

= 1 x 10^{-3} x 6.022 × 10^{23}

= 6.022 x 10^{20}

Ans. (a) Number of moles of carbon = 4.166 x 10^{-4} mol

(b) Number of carbon atoms = 6.022 x 10^{20}.

F. Arjun purchased 250 g of glucose (CHO) for Rs 40. Find the cost of glucose per mole.

Solution: Molar mass of glucose ($C_6H_{12}2O_6$)

= 6 x 12 + 12x1+6×16

= 180 g mol^{-1}

∵ 250 g of glucose costs 40

∴180 g of glucose will cost,

$\frac{40 \times 180}{250} = 28.8$

Ans. Cost of glucose per mole = ₹28.8.

G. The natural isotopic abundance of ^{10}B is 19.60% and ^{11}B is 80.40 %. The exact isotopic masses are 10.13 and 11.009 respectively. Calculate the average atomic mass of boron

Solution: Consider a boron sample containing 19.60% of ^{10}B and 80.40% of ^{11}B isotopes.
Average atomic mass of boron
$$= \frac{19.60 \times 10.013 + 80.40 \times 11.0009}{100}$$
$$= \frac{196.2548 + 885.1236}{100} = \frac{1081.3784}{100}$$
= 10.813784 ≅1081u
Ans. Average atomic mass of boron = 10.81 u [Note: In Textbook mass of ^{10}B is changed from 10.13 u 10.013 u]

H. Convert the following degree Celsius temperature to degree Fahrenheit.

a. 40 °C

b. 30 °C
Solution:
$$(a) \frac{°C}{5} = \frac{°F - 32}{9}$$
$$\therefore °F = \frac{9\,°C}{5} + 32$$
$$= \frac{9 \times 40}{4} + 32$$
=72 +32 = 104 °F

$$(b) °F = \frac{9\,°C}{5} + 32$$
$$= \frac{9 \times 30}{5} + 32$$
= 54 + 32 = 86 °F

Ans. (a) 104 °F
(b) 86 °F.

I. Calculate the number of moles and molecules of acetic acid present in 22 g of it.

Solution: Molar mass of acetic acid (CH_3COOH) = 60 g mol^{-1}
Number of moles of CH_3COOH = n
$$= \frac{W}{M} = \frac{22}{60}$$
= 0.3660 mol

Number of molecules of CH_3COOH
$= n \times N_A$
$= 0.3660 \times 6.022 \times 10^{23}$
$= 2.2076 \times 10^{23}$

Ans. Number of moles of CH_3COOH = 0.3660 mol Number of molecules of $CH_3COOH = 2.2076 \times 10^{23}$

J. 24 g of carbon reacts with some oxygen to make 88 grams of carbon dioxide. Find out how much oxygen must have been used.

Solution: Mass of carbon + Mass of oxygen-Mass of CO_2
∴ Mass of oxygen
= Mass of CO_2 - Mass of carbon
=88 - 24 = 64 g
Ans. Mass of oxygen used = 64 g.

K. Calculate number of atoms is each of the following. (Average atomic mass: N = 14 u, S = 32 u)
a. 0.4 mole of nitrogen
b. 1.6 g of sulfur

Solution: (a) Number of molecules of $N_2 = 0.4 \times 6.022 \times 10^{23}$
∴ Number of N atoms $= 0.4 \times 6.022 \times 10^{23} \times 2 = 4.8176 \times 10^{23}$

(b) Number of moles of S
$= \frac{W}{M} = \frac{1.6}{32} mol$
Number of atoms of S
$= \frac{1.6}{32} \times 6.022 \times 10^{23}$
$= 3.011 \times 10^{23}$

Ans. (a) Number of N atoms$=4.8176 \times 10^{23}$

(b) Number of atoms of S$= 3.011 \times 10^{22}$

L. 2.0 g of a metal burnt in oxygen gave 3.2 g of its oxide. 1.42 g of the same metal heated in steam gave 2.27 of its oxide. Which law is verified by these data?

Solution: Mass of oxygen in metal oxide = 3.2-2=1.2 g Mass of oxygen in second case = 2.27 -1.42 = 0.85 g
∵ 0.85 g oxygen combines with 1.42 g metal
∴1.2 g oxygen will cambine with,
$\frac{1.42 \times 1.2}{0.85} = 2.0 g\ metal.$
This observation represents a law of definite proportion.

M. In two moles of acetaldehyde (CH_3CHO) calculate the following

a. Number of moles of carbon

b. Number of moles of hydrogen

c. Number of moles of oxygen

d. Number of molecules of acetaldehyde

Solution: (a) ∵ 1 mol of acetaldehyde, CH_3CHO contains 2 mol of carbon
∴ 2 mol of acetaldehyde contains 4 mol carbon.

(b)∵ 1 mol CH_3CHO contains 4 mol H
∴ 2 mol CH_3CHO contain 2 x 4 = 8 mol H

(c)∵ 1 mol CH_3CHO contains 1 mol oxygen
∴2 mol CH_3CHO contain 2 mol oxygen

(d) Number of molecules of $CH_3CHO = 2 \times 6.022 \times 10^{23}$
$= 1.2044 \times 10^{24}$ molecules

Ans. (a) 4 mol C

(b) 8 mol H

(c) 1 mol oxygen

(d) 1.2044×102^{24} molecules.

N. Calculate the number of moles of magnesium oxide, MgO in i. 80 g and ii. 10 g of the compound. (Average atomic masses of Mg = 24 and O=16)
Solution: (i) Molar mass of MgO=24+16=40 g mol^{-1}
∴ Number of moles of MgO = n

$$= \frac{W}{M}$$

$$= \frac{80}{40} = 2\ mol$$

(ii) Number of moles of MgO

$$= \frac{10}{40} = 0.25 \text{ mol.}$$

Ans. (1) Number of moles of MgO = 2 mol
(2) Number of moles of MgO = 0.25 Mol.

O. What is volume of carbon dioxide, CO_2 occupying by i. 5 moles and ii. 0.5 mole of CO_2 gas measured at STP.

Solution: (i) ∵ 1 mol CO_2 at STP occupies 22.4 dm^3
∴ 5 mol CO_2 at STP occupies, 22.4x5=112 dm^2

(ii) Volume of CO_2 at STP = 0.5 x 22.4=11.2 dm^3

Ans. (i) Volume of CO_2 = 112 dm^3
(ii) Volume of CO_2 = 11.2 dm^3

P. Calculate the mass of potassium chlorate required to liberate 6.72 dm^3 of oxygen at STP. Molar mass of $KClO_3$ is 122.5 g mol^{-1}.

Solution:
$2KClO_3 \rightarrow 2\ KCl + 3O_2$
2 mol 3 mol

Volume of 3 mol O_2 at STP = 3 x 22.4 = 67.2 dm^3
∵ 67.2 dm^3 of O_2 is liberated at STP by 2 mol $KClO_3$
∴ 6.72 dm^3 of O_2 at STP will be liberated by,
$\frac{6.72}{67.2} \times 2 = 0.2 \text{ mol } KClO_3$
∴Mass of $KClO_3$ required = 0.2 x 122.5 = 24.5 g

Ans. Mass of $KClO_3$ required = 24.5 g.

Q. Calculate the number of atoms of hydrogen present in 5.6 g of urea, $(NH_2)_2CO$. Also calculate the number of atoms of N, C and O.
Solution: Molar mass of urea, $(NH_2)_2$ CO=60g mol^{-1}
Number of moles of urea

$= \frac{W}{M} = \frac{5.6}{60} = 0.09333$ and
∴Number of molecules of urea =
n x N_A

= 0.09333 x 6.022 × 10^{23}

= 0.5620 x 10^{23}
∴ Number of H atoms = 4 x 0.5620 x 10^{23} = 2.248 x 10^{23}
Number of N atoms = 2 x 0.5620 × 10^{23}

= 1.124 x 10^{23}
Number of C atoms=5.620 × 10^{22}

Number of O atoms $= 5.620 \times 10^{22}$

Ans. Number of H atoms $= 2.248 \times 10^{23}$

Number of N atoms $= 1.124 \times 10^{23}$

Number of C atoms$=5.620 \times 10^{22}$

Number of O atoms $= 5.620 \times 10^{22}$

R. Calculate the mass of sulfur dioxide produced by burning 16 g of sulfur in excess of oxygen in contact process. (Average atomic mass : S = 32 u, O=16 u)
Solution:
$S \quad + \quad \rightarrow \quad SO_2$
32g 1 mole

$\because$ 32 g sulphur produces 1 mole SO_2

$\therefore$16 g sulphur will produce 0.5 mol SO_2.
Molar mass of $SO_2 = 64$ g mol^{-1}

$\therefore$ Mass of SO_2 produced $= 0.5 \times 64 = 32$ g

Ans. Mass of SO_2 produced = 32 g.

5. Explain

A. The need of the term average atomic mass.

Ans. An element may have different isotopes i.e., atoms having same number of electrons and protons but different number of neutrons. The isotopes have different atomic masses and they exist in different proportion or abundance. The atomic mass of an element is the weighted average of atomic masses of its isotopes. This is called average atomic mass. Isotopes with relative abundances and atomic masses as shown against each of them.

Isotope	Atomic mass (u)	Relative Abundance (%)
^{12}C	12.00000	98.892
^{13}C	13.00335	1.108
^{14}C	14.00317	2×10^{-10}

From the above data, the average atomic mass of carbon

$=(12 \text{ u}) (98.892 / 100) +(13.00335 \text{ u})$

(1.108/100)+(14.00317) (2×10^{-10}/100)

= 12.011 u
Similarly, average atomic masses for other elements can be calculated.

B. Molar mass.

Ans. Molar Mass: The mass of one mole of a substance (element/compound) in grams is called its molar mass. The molar mass of any element in grams is numerically equal to atomic mass of that element in u.

Explanation:

Element	Atomic mass (u)	Molar mass (g mol^{-1})
H	1.0 u	1.0g mol^{-1}
C	12.0u	12.0g^{-1}
O	16.0u	16.0g^{-1}

Similary molar mass of any substance, existing as polyatomic molecule, in grams is numerically equal to its molecular mass or formula mass in u.

Polyatomic substance	Molecular formula mass (u)	Molar mass (g mol^{-1})
O_2	32.0 u	32.0 g mol^{-1}
H_2O	18.0 u	18.0 g mol^{-1}
NaCl	58.5 u	58.5 g mol^{-1}

C. Mole concept.

Ans. Mole: One mole of a substance is defined as the amount of a substance that contains the number of particles, atoms, molecules, ions or electrons equal to the number of carbon atoms, etc. present in 0.012 kg of Carbon-12 i.e., 6.0221367×10^{23} particles.
One mole of a substance contains 6.022×10^{23} molecules while one mole (or one gram atom) of an element contains 6.0221367×10^{23} atoms.

[Note: The name of the unit is mole and the symbol for the unit is mol.]

D. Formula mass with an example.

Ans. i. The formula mass of a substance is the sum of atomic masses of the atoms present in the formula.
ii. In substances such as sodium chloride, positive (sodium), and negative (chloride) entities are arranged in a three dimensional structure in a way that one sodium (Na^+) ion is surrounded by six chlorides (Cl^-) ions, all at the same distance from it and vice versa. Thus, sodium chloride does not contain discrete

molecules as the constituent units.
iii. Therefore, NaCl is just the formula that is used to represent sodium chloride though it is not a molecule. iv. In such compounds, the formula (i.e., NaCl) is used to calculate the formula mass instead of molecular mass.
e.g. Formula mass of sodium chloride = atomic mass of sodium + atomic mass of chlorine atomic mass of chlorine
= 23,0 U + 35.5 u = 58.5 u

E. Molar volume of gas.

Ans. The volume of one mole of a gas is called molar volume. At STP condition (0 °C and 1 atm), the molar volume of any gas is 22.4 dm^3.

F. Types of matter (on the basis of chemical composition).

Ans. Matter on the basis of chemical composition can be classified as follows:
i. Pure substances: They always have a definite chemical composition. They always have the same properties regardless of their origin. e.g. Pure metal, distilled water, etc.
They are of two types:

a. Elements: They are pure substances, which cannot be broken down into simpler substances by ordinary chemical changes.
Elements are further classified into three types:

1. Metals: i. They have a lustre (a shiny appearance).

ii. They conduct heat and electricity.

iii. They can be drawn into wire (ductile).

iv. They can be hammered into thin sheets (malleable).
e.g. Gold, silver, copper, iron. Mercury is a liquid metal at room temperature.

2. Nonmetals: i. They have no lustre. (except diamond, iodine)

ii. They are poor conductors of heat and electricity. (except graphite) iii. They cannot be hammered into sheets or drawn into wire, because they are brittle.
e.g. lodine

3. Metalloids: Some elements have properties that are intermediate between metals and nonmetals and are called metalloids or semimetals. e.g. Arsenic, silicon, and germanium.

b. Compounds: They are the pure substances that are made up of two or more elements in a fixed proportion. e.g. Water, ammonia, methane, etc

ii. Mixtures: They have no definite chemical composition and hence no definite properties. They can be separated by physical methods. e.g. Paint (a mixture of oils, pigment, additive), concrete (a mixture of sand, cement, water), etc.

a. Mixtures are of two types: a. Homogeneous mixture: In homogeneous mixture, constituents remain uniformly mixed throughout its bulk. e.g. Solution, in which solute and solvent molecules are uniformly mixed throughout its bulk.

b. Heterogeneous mixture: In heterogeneous mixture, constituents are not uniformly mixed throughout its bulk. e.g. Suspension, which contains insoluble solid in a liquid.