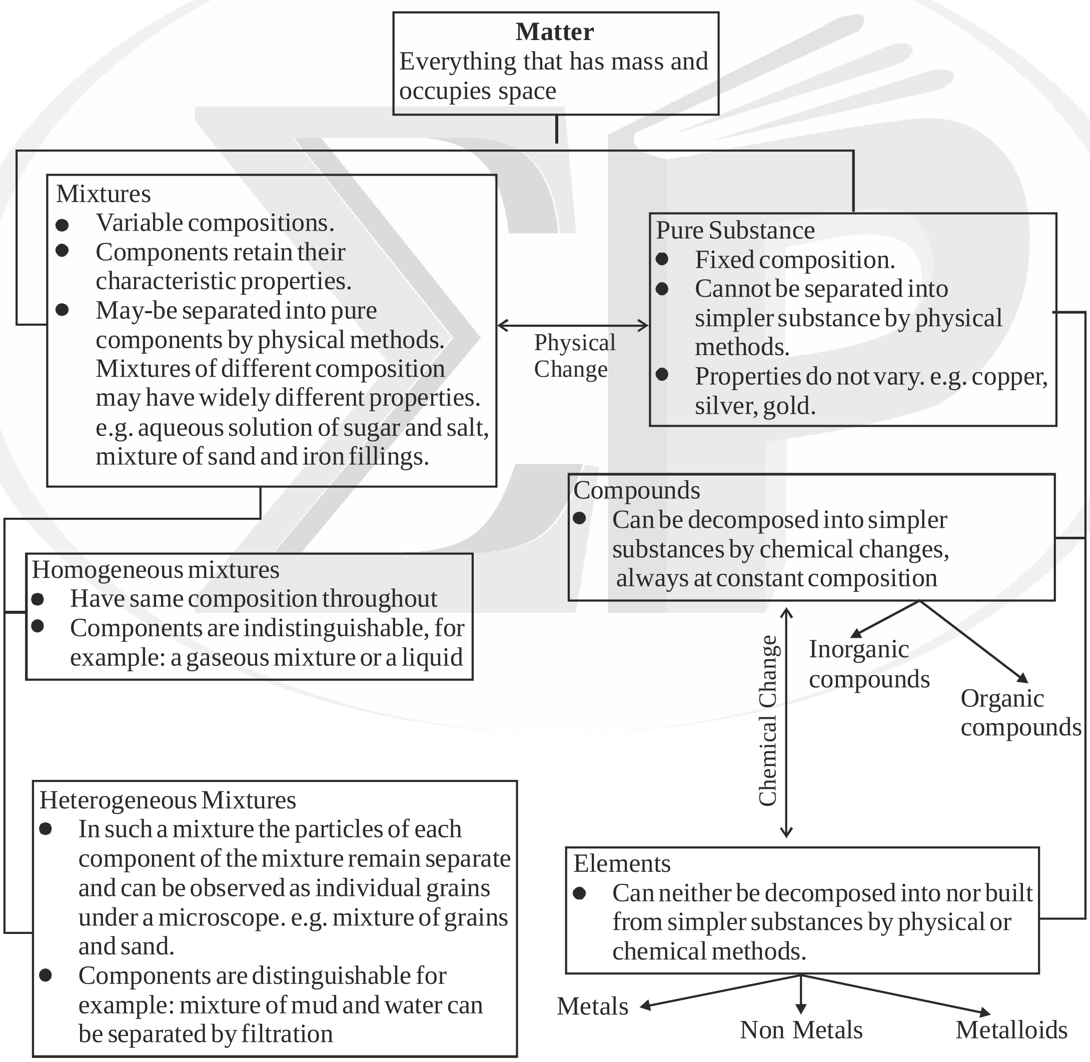
XI CHEMISTRY NOTES ON SOME BASIC CONCEPTS OF CHEMISTRY

GENERAL INTRODUCTION AND NATURE OF MATTER

- Anything that occupies space is called matter
- All substances contain matter which can exist in three states - solid, liquid and gas.
- A solid has definite shape and a definite volume.
- A liquid has definite volume but not definite shape.
- A gas neither has a definite volume nor a definite shape.
- These three forms of matter are interconvertible by changing the conditions of temperature and pressure.

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

Matter can also be classified into elements, compounds or mixtures.



- **Unit and Measurement:** Out of many systems of measurements SI (International System of Units) is commonly used.

Example: length is measured in 'm'; mass in kg etc. This was done by a study committee of the French Academy of science. Later on it was adopted by general conference of weights and measures and termed International System of Units or SI.

- The SI system has seven base units which are listed as

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

- **The definitions of SI base units of some important quantities are given in following table.**

Measured Quantity	Unit	Definition
Length	metre	The metre is the length of path travelled by light in vacum during a time interval of 1/299, 792, 458 of a second.
Mass	kilogram	A kilogram is equal to the mass of the international prototype of the kilogram. It is also defined as the mass of platinum bock store at the Internal Bureau of weights and measures in France.
Time	second	The second is the duration of 9,192,631,770 periods of the radiation corresponding to the transition between the two hyperfine levels of the ground state of the cesium-133 atom.
Electric current	ampere	The ampere is the constant current which, if maintained in two straight parallel conductors of infinite length, of negligible cross section, and placed one metre apart in vacuum, would produce between these conductors a force equal to 2×10^{-7} newton per meter of length.
Thermodynamic temperature	kelvin	A kelvin is equals to exactly 1/273.16 of the thermodynamic temperature of the triple point of water.
Amount of substance	mole	Mole is the amount of substance of a system which contains as many elementary entities as there are atoms in 0.012 kilogram of carbon –12.
Luminous intensity	candela	The candela is the luminous intensity, in a given direction, of a source that emits monochromatic radiation of frequency 540×10^{12} hertz and that has a radiant intensity in that direction of 1/683 watt per steradian.

- **Prefixes:-** The SI units of some of the physical quantities are too small or too large. To change the order of magnitude, these are expressed by using prefixes before the name of the base units.
- **Some commonly used prefixes for reducing or enlarging the size of any unit.**

Symbol	Prefix	Multiplying Factor	Symbol	Prefix	Multiplying Factor
d	deci	10^{-1}	da	deca	10^{1}
c	centi	10^{-2}	k	kilo	10^{3}
m	milli	10^{-3}	M	mega	10^{6}
μ	micro	10^{-6}	G	giga	10^{9}
n	nano	10^{-9}	T	tera	10^{12}
p	pico	10^{-12}	E	exa	10^{18}

- **Derived Units:-** There are many physical quantities which are measured in the laboratory. Their units are obtained by using SI base units. For example to determine the volume of box of sides 1.0m, 2.0m, and 2.5m, we multiply its height, breadth and length to get the volume $1 \times 2 \times 2.5 = 5.0\ m^3$. Hence '$m^3$' is the derived unit of volume in SI unit.
- **Some common derived units.**

	Physical Quantity	Definition	Units	Symbol
1.	Area	Length square	Square metre	m^2
2.	Volume	Length cube	Cubic metre	m^3
3.	Density	Mass/unit vol.	Kilogram per cubic metre	$kg\ m^{-3}$
4.	Force	Mass × acceleration	Newton	$N = kg\ ms^{-2}$
5.	Pressure	Force/unit area	Pascal (Newton per sq. metre)	$Pa = Nm^{-2} = kgm^{-1}s^{-2}$
6.	Velocity	Distance/unit time	Metre per second	ms^{-1}
7.	Acceleration	Speed change/unit time	Metre per second/second	ms^{-2}
8.	Electric charge	Current × time	Coulomb	$C = As$
9.	Potential difference	—	Volt	$V = kg\ m^2s^{-3}\ A^{-1}$ $= JA^{-1}\ s^{-1} = JC^{-1}$
10.	Electric resistance	Pot. diff/current	ohm	$\Omega = VA^{-1}$
11.	Electric conductance	Reciprocal of resistance	ohm^{-1}	$\Omega^{-1} = AV^{-1}$
12.	Frequency	Cycles/sec.	Hertz	$Hz = s^{-1}$
13.	Work, energy	Force × Distance	Joule	$J = Nm = kg\ m^2\ s^{-2}$

- **The definitions of SI base units of some important quantities are given in following table.**
- **Accuracy:-**The accuracy is a measure of the difference between the experimental value or the mean value of a set of measurement and the true value i.e. Accuracy = Mean value – True value.
- **Precision:-** It is denoted as the difference between a measured value and the arithmetic mean value for series of measurements.
- Precision - Individual value-Arithmetic mean value
- **Exponential notation:-** The exponential notations (in which any number can be represented in the form $N \times 10^n$ where n is any positive or negative integer) are used in expressing very small and very large numbers.
- Numbers can be expressed as a unit number (a number between one and ten which may be a decimal, e.g. 7.32) times some power of ten(i.e., in scientific notation). The exponent of ten is determined by number of places the decimal point must be moved to form the unit number. For every place the decimal is moved to the left, the exponent is increased by one unit and when decimal is moved to one place right, the exponent is reduced by one unit. For example: $0.00378 = 3.78 \times 10^{-3}$; $651 = 6.51 \times 10^2$; $0.087 \times 10^8 = 8.7 \times 10^6$.
- **Significant figures:-** Significant figures in a number are all certain digits plus one doubtful digit. For example a figure 10.52 has four significant figures, the last digit (2) having a greater uncertainty than other digits. Note that the decimal point does not determine the number of significant figures. There are several general rules governing the significance of zeros: a) Final zeros after a decimal points are significant figures; b) Final zeroes before a decimal points may or may not be significant. c) If a number is less than one, zeros following the decimal point are not significant but those following the digits would be, e.g., <u>0.00</u>12300, the zeros underlined are not significant but the zeros after 123 are significant.
- **Rounding off:-** It means dropping of uncertain digits in a number. Following are the points to remember:-

(i) If the right most digit to be removed is more than 5, the preceding number is increased by one. For example :- 3.19 is rounded off to 3.2

(ii) If the right most digit to be removed is less than 5, the preceding number remain unchanged e.g., 3.73 is rounded off to 3.7

(iii) If the right most digit to be removed is equal to 5, the preceding number is retained as such and not changed if it is even, but increased by one if it is odd. e.g., 3.75 is rounded off to 3.8 and 3.65 is rounded off to 3.6.

- **Dimensional analysis:-** This is a method to convert unit from one system to other.
- **Some important conversion factors**

1 mile	=	1760 yards	1 metric ton	=	1000 kg
1 yard	=	3 feet	1 kg	=	1000 g
1 foot	=	12 inches	1 g	=	1000 mg
1 inch	=	2.54 cm	1 lb	=	454.0 g
	=	1000 ml		=	0.454 kg
	=	1000 cm^3 = 10^3cc = 1 dm^3			

- **From given unit to S.I. unit**

1 Å = 10^{-10}m = 10^{-8} cm			1 atm	=	760mm or torr
				=	101,325 Pa or Nm^{-2}
				=	1.013×10^6 dynes/cm^2
1 a.m.u.	=	1.66053×10^{-27} kg	1 bar	=	10^5 Nm^{-2} = 10^5 Pa
t°C	=	t + 273 K	1 mm or torr	=	133.322 Pa or Nm^{-2}
1 dyne	=	10^{-5}N	1 electron volt (eV) = 1.6022×10^{-19}J.		
1 calorie	=	4.184J ; 1 erg = 10^{-7} J			

ILLUSTRATIONS

1. Express the following numbers upto three significant figures.

a. 306.35 **b.** 0.0038816 **c.** 178975×10^4 **d.** 0.25400 **e.** 2.65986×10^3

Sol. **a.** 306.35 = 306 **b.** 0.0038816 = 3.88×10^{-3}

c. $1.78975 \times 10^4 = 1.79 \times 10^4$ [rounded off according to rule vi (b)]

d. $0.25400 = 2.54 \times 10^{-1}$ **e.** $0.25986 \times 10^3 = 2.66 \times 10^3$ [rounded off according to rule vi (b)]

2. Add the following quantities expressed in grams, to the proper number of significant figures.

a.	**b.**	**c.**	**d.**
23.540	58.0	4.20	415.5
5.465	0.0038	1.6523	3.64
0.322	0.00001	0.015	0.238

Sol. **a.**
25.340
5.465
+ 0.322 (all have 3 decimal places)
31.127 g (answer should be reported to 3 decimal places)

Correct answer = 31.127 g

b.
58.0 (has only one decimal place)
0.0038
+ 0.00001
58.00381 g (answer should be reported to one decimal place)

Correct answer = 58.0 g

c.
4.20 (has two decimal places)
1.6523
+0.015
5.8673 g (answer should be reported to two decimal places after rounding off)

Correct answer = 5.87 g

$$\begin{array}{ll} \ \ 415.5 & \text{(has only decimal place)} \\ \ \ \ \ 3.64 & \\ \textbf{d.}\ \ + 0.238 & \\ \hline \ \ 419.378 \text{ g} & \text{(answer should be reported to one decimal place after rounding off)} \\ \hline \end{array}$$

Correct answer = 419.4 g

3. Calculate to the proper number of signifcant digits :

a. $(3.50) \times 10^2$ m) + $(3.00 \times 10^6$ mm) **b.** $(2.50 \times 10^2$ cm) $(2.00 \times 10^8$ cm)

Sol. **a.** $(3.50) \times 10^2$ m) + $(3.00 \times 10^6$ mm) or $(0.35 \times 10^3$ m) + $(3.00 \times 10^3$ m) or 3.35×10^3 m

b. $(2.50 \times 10^2$ cm) $(2.00 \times 10^8$ cm) or 5.00×10^8 cm^2

4. The average speed of a bus is 95 mile per hour. Express the speed in metre per second. Given 1 mile = 1.6 km.

Sol. Given 1 mile = 1.6 km = 1.6×10^3 m

$$\text{Conversion factor} = \frac{1.6\times10^3 \text{ m}}{1 \text{ mile}}; \quad 1 \text{ hr} = 60 \times 60\text{s} = 3600 \text{ s}$$

$$\text{Conversion factor} = \frac{3600 \text{ s}}{1 \text{ hr}}; \quad \text{Speed} = \frac{95 \text{ mile}}{1 \text{ hr}} = \frac{95 \text{ mile}}{1 \text{ hr}} \times \frac{1.6\times10^3 \text{ m}}{1 \text{ mile}} \times \frac{1 \text{ hr}}{3600\text{s}} = 42.2 \text{ ms}^{-1}$$

5. A wood block, 10 inch × 6.0 inch × 2.0 inch, weighs 3 lb 10 oz.Calculate its density in (i) g/cm^3 and (ii) kg/m^3.

Sol. Volume = 10 inch × 6.0 inch ×2.0 inch = 120 inch^3

$$\text{We know that } 2.54 \text{ cm} = 1 \text{ inch} \quad ; \quad \text{Conversion factor} = \frac{2.54 \text{ cm}}{1 \text{ inch}}$$

$$\text{Hence, } 120 \text{ inch}^3 = (12\text{- inch}^3) \times \left(\frac{2.54 \text{ cm}}{1 \text{ inch}}\right)^3 = 1966 \text{ cm}^3$$

$$\text{Mass} = 58 \text{ oz} \quad ; \text{ We know that } 1 \text{ oz} = 58 \text{ oz} \times \frac{28.35 \text{ g}}{1 \text{ oz}} = 1644.3 \text{ g}$$

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}} = \frac{1644.3 \text{ g}}{1996 \text{ cm}^3} = 0.836 \text{ g/cm}^3 = \frac{0.836 \text{ g}}{\text{cm}^3} \times \frac{1\text{kg}}{1000 \text{ g}} \times \left(\frac{100 \text{ cm}}{1\text{m}}\right)^3 = 836 \text{ kg/m}^3.$$

6. Determine the mass in kg of 20.0 ft^3 of aluminium. The density of aluminium is 2.70 g cm^{-3}.

Sol. 1 ft = 12 × 2.54 cm ; $\text{Conversion factor} = \dfrac{12\times2.54 \text{ cm}}{1 \text{ ft}}$

$$\text{Volume} = 20.0 \text{ ft}^3 = 20.0 \text{ ft}^3 \times \left(\frac{12\times2.54 \text{ cm}}{1 \text{ ft}}\right)^3 = 20 \times (12 \times 2.54 \text{ cm})^3$$

$$\text{Mass} = \text{volume} \times \text{density} = 20 \times (12 \times 2.54 \text{ cm})^3 \times \frac{2.70\text{g}}{\text{cm}^3} = 15.92 \times 10^5 \text{ g} = 15.29 \times 10^2 \text{ kg}.$$

PRACTICE PROBLEMS

1. How many seconds are there in one year (365 days)?

2. Covert5 the following into kilograms : **a.** 500 Mg (mass of ship) **b.** 1 fg (mass of RNA molecule)

3. A tennis ball was observed to travel at a speed of 98 mile per hour. Express the speed in SI unit.

4. Convert the following into metre : **a.** 41 pm **b.** 40 Em **c.** 1.5 gm

5. Calculate the total mass of ring in grams which contains 0.5 carat diamond and 7.00 gram gold.
[Given : 1 carat = 3.168 grains ; 1 gram = 15.4 grains]

6. Convert 40 calories into joules.

7. Density of an object is 9.4 g cm^{-3}. Express the density in kg m^{-3}.

TRY IT YOURSELF

1. Classify the following into metals and non-metals

   i. Helium **ii.** Sodium **iii.** Mercury **iv.** Graphite

   vi. Carbon **vii.** Lead **viii.** Magnesium **ix.** Chlorine **x.** Phosphorus.

2. Why is mixture of salt in water a solution while that of oil and water is not?

3. Classify the following into elements, compounds and mixtures.

   i. Marble **ii.** Honey **iii.** Tooth paste **iv.** Sugar **v.** Gold

   vi. Brass **vii.** Nitre **viii.** Nitrogen **ix.** Slaked lime **x.** Fruit juice.

4. Give the names of two elements which acts as metalloids.

5. Name a mixture used **: i.** by all living beings **ii.** in the construction of building **iii.** as a food.

LAWS OF CHEMICAL COMBINATION

Law of Conservation of Mass
This law is given by Lavoisier. Matter can neither be created nor be destroyed.

Law of Definite Proportion (Louis Proust)
A sample of pure compound whatever is its source, always consist of same elements combined in the same proportion by mass.
e.g. Pure water from any source say rain, well or lakes would contain 2 atoms of hydrogen and 1 atom of oxygen.

Law of Multiple Proportions
When two elements combine to give two or more compounds, the mass of element which combines with the fixed mass of other element bear a simple ratio.
e.g. 'C' and 'O' combine to give CO and CO_2, thus the mass of 'O' which combine with fixed mass of "C"(12g) in CO and CO_2 bears a simple ratio of 16:32 or 1:2
This law was formulated by Dalton.

Gay-Lussac's Law
This law states that when gases react with each other, their volumes bear a simple whole no. ratio to one another and to volume of products (if gases) and similar conditions of pressure and temperature.

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

On the basis of laws of chemical combination Dalton stated that all matter is made up of atoms which are indivisible and indestructible ultimate particles. All the atoms of a given element are identical, both in mass and chemical properties, however atoms of different elements have different masses and different chemical properties. Atoms combine with one another to form molecules and atoms are smallest particle that takes part in a chemical reaction. These points are summerised as Dalton's atomic theory. (The atomic mass of an element is expressed relative to 12C isotope of carbon which has an exact value of 12).

ILLUSTRATIONS

1. What mass of sodium chloride would be decomposed by 9.8 g of sulphuric acid if 12 g of sodium bisulphate and 2.75 g of hydrogen chloride were produced in a reaction assuming that the law of conservation of mass is true?

Sol. $NaCl + H_2SO_4 = NaHSO_4 + HCl$

According to law of conservation of mass,

Total masses of reactants = Total masses of products

Let the mass of NaCl decomposed be x g ; so $x + 9.8 = 12.0 + 2.75 = 14.75$; $x = 4.95$ g

2. In an experiment, 2.4 g of iron oxide on reduction with hydrogen yield 1.68 g of iron. In another experiment 2.9 g of iron oxide give 2.03 g of iron on reduction with hydrogen. Show that the above data illustrate the law of constant proportions.

Sol. **In the first experiment**

The mass of iron oxide = 2.4 g ; The mass of iron after reduction = 1.68 g

The mass of oxygen = Mass of iron oxide – Mass of iron = (2.4 –1.68) = 0.72 g

Ratio of oxygen and iron = 0.721 : 1.68 = 1 : 2.33

In the second experiment

The mass of iron oxide = 2.9 g ; The mass of iron after reduction = 2.03 g

The mass of oxygen = (2.9 –2.03) = 0.87 g ; Ratio of oxygen and iron = 0.87 : 2.03 = 1 : 2.33

3. Carbon combines with hydrogen to form three compounds A, B and C. The percentages of hydrogen in A, B and C are 25, 14.3 and 7.7 respectively. Which law of chemical combination is illustrated?

Sol.

Compound	% of hydrogen	% of carbon
A	25.0	(100 –25.0) = 75.0
B	14.3	(100 –14.3) = 85.7
C	**7.7**	(100 –7.7) = 92.3

In Compound A

25 parts of hydrogen combine with 75 parts of carbon

1 part of hydrogen combines with 75/25 = 3 parts of carbon

In Compound B

14.3 parts of hydrogen combine with 85.7 parts of carbon

1 part of hydrogen combines with 85.7/14.3 = 6.0 parts of carbon

In Compound C

7.7 parts of hydrogen combine with 92.3 parts of carbon

1 part of hydrogen combines with 92.3/7.7 = 12.0 parts of carbon

Thus, the masses of carbon in three compounds A, B and C, which combine with a fixed mass of hydrogen are in the ratio of 3 : 6 : 12 or 1 : 2 : 4. This is a simple ratio. Hennce, the data illustrate the law of multiple proportions.

4. Two compounds each containing only tin and oxygen had the following composition :

	Mass % of tin	Mass % of oxygen
Compound A	78.77	21.23
Compound B	88.12	11.88

Show how data illustrate the law of multiple proportions?

Sol. **In Compound A**

21.23 parts of oxygen combine with78.77 parts of tin 1 part of oxygen combines with 78.77/21.23 = 3.7 parts of tin

In Compound B

11.88 parts of oxygen combine iwth 88.12 parts of tin

1 part of oxygen combines with 88.12/11.88 = 7.4 parts of tin

Thus, the mass of tin in compound A and B which combine with a fixed mass of oxygen are in the ratio of 3.7 : 7.4 or 1 : 2. This is a simple ratio.Hence, the data illustrate the law of multiple proportions.

5. How much volume of oxygen will be required for complete combustion of 40 mL of acetylene (C_2H_2) and how much volume of carbon dioxide will be formed? All volumes are measured at NTP.

Sol.

$$\begin{array}{llll} 2C_2H_2 + & 5O_2 & = 4CO_2 & + \; 2H_2O \\ \text{2 vol.} & \text{5 vol.} & \text{4 vol.} & \\ \text{40 mL} & \frac{5}{2}\times 40\text{ mL} & \frac{4}{2}\times 40\text{ mL} & \\ \text{40 mL} & \text{100 mL} & \text{80 mL} & \end{array}$$

So, for complete combustion of 40 mL of acetylene, 100 mL of oxygen are required and 80 mL of carbon dioxide are formed.

6. Show that the following results illustrate the law of reciprocal proportions :

a. 0.92 g Mg produces 1.54 g magnesium oxide.

b. 0.41 g Mg liberates 380 mL of hydrogen at NTP from an acid.

c. 1.25 g of water results from the union of 1.11 g of oxygen with hydrogen. (Denisty of hydrogen at NTP = 0.00009 g mL^{-1}).

Sol. **a.** Mass of oxygen = (1.54 –0.92) = 0.62 g ; 0.92 g Mg combines with oxygen = 0.62 g

1 g Mg combines with oxygen $= \dfrac{0.62}{0.92} = 0.6739$ g

b. Mass of 380 mL of hydrogen at NTP = 380 × 0.00009 g = 0.0342 g

1 g Mg liberates hydrogen $= \dfrac{0.0342}{0.41} = 0.0834\ g$

The ratio of hydrogen and oxygen which combines with 1 g Mg is 0.0834 : 06739, i.e., 1 : 8

c. Mass of hydrogen = (1.25 –1.11) = 0.14 g

The ratio of hydrogen and oxygen in water = 0.14 : 1.11 or 1 : 8.

Since, the ratio of hydrogen and oxygen is same,the law of reciprocal proportions is illustrated.

PRACTICE PROBLEMS

1. If 6.3 g $NaHCO_3$ are added to 15 g of CH_3COOH solution, the residue is found to weigh 18 g. What is the mass of CO_2 released in the reaction?

2. *x* g of potassium chlorate on decomposition produced 1.92 g of oxygen and 2.96 g of potassium chloride. What is the value of *x*?

3. What weight of sodium chloride is decomposed by 4.9 g of sulphuric acid, if 6 g of sodium bisulphate ($NaHSO_4$) and 1.825 g of hydrogen chloride are produced in the same reaction?

4. If the law of constant composition is true, what weights of calcium, carbon and oxygen are present in 1.5 g calcium carbonate? Given that the sample of calcium carbonate contains 40% Ca; 12% Carbon; 48% Oxygen.

5. Two oxides of a metal contain 27.6% and 30% oxygen respectively. If the formula of first oxide is M_3O_4, find that of the second.

6. A metal forms two osixes, one contains 46.67% of the metal and another 63.94% of the metal. Show that these results are in accordance with the law of multiple proportion.

7. Aluminium carbide contains 75% of Al; aluminium oxide contains 52.9% of Al and carbon dioxide contains 27.27%C. Show that these figures illustrate the law of reciprocal proportions.

TRY IT YOURSELF

1. If ten volume of dihydrogen gas react with five volume of dioxygen gas, how many volume of water vapour would be produced?
2. 1.375 g of cupric oxide was reduced by heating in a current of hydrogen and the weight of copper that remained was 1.098 g. In another experiment, 1.179 g of copper was dissolved in the nitric acid and the resulting copper nitrate was converted into cupric oxide by ignition. The weight of cupric oxide formed was 1.476 g. Show that these results illustrate the law of constant composition.
3. 5.975 g of the higher oxide of metal gave 5.575 g of lower oxide on heating. The quantity of the lower oxide gave 5.175 g of metal on reduction. Prove that these results are in accordance with the law of multiple proportions.
4. Air contains 20% oxygen by volume. Calculate the theoretical volume of air which will be required for burning completely 500 m^3 of acetylene gas. All volumes are measured under the same conditions of temperature and pressure.
5. 10 mL of H_2 combine with 5 mL of O_2 to form water. When 200 mL of H_2 at STP is passed over heated CuO, the CuO losses 0.144 g of its weight. Does the above data correspond to the law of constant composition?
6. A box contains some identical red coloured balls, labelled as A, each weighing 2 g. Another box contains identical blue coloured balls, labelled as B_1 each weighing 5 g. Consider the combinations AB, AB_2, A_2B and A_2BN_3 and show that law of multiple proportion is applicable.
7. 45.4 L of dinitrogen reacted with 22.7 L of dioxygen and 45.4 L nitrous oxide was formed.

   The reaction is given below : $2N_2(g) + O_2(g) \longrightarrow 2N_2O(g)$

   Which law is being obeyed in this experiment? Write the statement of the law.
8. Describe what you need to do in the laboratory to test

   (i) the law of conservation of mass. (ii) the law of definite proportion and (iii) the law of multiple proportions.

DIFFERENT MASSES, MOLE CONCEPTS AND FORMULAE OF COMPOUNDS

- The ***'average atomic mass'*** is obtained by taking into account the natural abundance of different isotopes of that element.
- Atomic mass = $\dfrac{\text{Average mass of on atom}}{1/12 \times \text{mass of an atom of C}-12}$
- Average Atomic mass = $\dfrac{\text{R.A}(1) \times \text{mass number} + \text{R.A.}(2) \times \text{mass number}}{\text{R.A.}(1) + \text{R.A.}(2)}$
- The ***'molecular mass'*** of a molecule is obtained by taking sum of the atomic masses of different atoms present in the molecule.

  Mass of CO_2 = Mass of C atom + 2 Mass of O atom = 12 + 2 × 16 = 44
- Average molecular mass = Σ fractional abundance × molar mass ; Molecular mass = 2 × V.D.
- The ***'formula mass'*** of a substance (ionic substances) is the sum of the atomic masses of all the atoms in a formula unit of a compound. e.g., Formula mass of NaCl = 23 + 35.5 = 58.5 u).
- **Percentage of composition:-** Mass percentage composition of a compound gives the mass of each element expressed as the percentage of the total mass.
- **Molecular formula:-** The molecular formula can be calculated by determining the mass percent of different elements presents in the compound and its molecular mass. It is the actual formula of a compound showing the total number of atoms of constituent elements e.g. C_6H_6 is molecular formula of benzene.
- **Empirical formula :-** The relative number of each kind of atom in a molecule may be calculated from the mass of the various atoms in the compound. The formula which gives the simplest ratio of numbers of each kind of atom in

a compound is called the empirical formula and may differ from molecular formula. For example, benzene has a molecular formula of C_6H_6 and an empirical formula of CH.

Molecular formula = n × empirical formula, when n is simple

- **Avagadro's number :-** One mole of any element contains 6.023×10^{23} atoms of that element. This number is known as Avogadro's number.
- Volume occupied by one mole of a gaseous substance is called ***gram molecular volume.*** Its value is 22.4 Litre at N.T.P.
- The mass of one mole of a substance in grams is called its ***'molar mass'.***
- Number of moles = $\dfrac{\text{Given weight}}{\text{At. wt/mol. wt}} = \dfrac{\text{number of atoms/molecules/electrons}}{6.023 \times 10^{23}}$ = 22.4 litres at STP (T = 273 K; P = 1 atm)

ILLUSTRATIONS

1. Calculate the mass of 1.5 gram molecule of sulphuric acid.

Sol. Molecular mass of H_2SO_4 = $2 \times 1 + 32 + 4 \times 16$ = 98.0 amu

Gram-molecular mass of H_2SO_4 = 98.0 g; Mass of 1.5 gram molecule of H_2SO_4 = 98.0×1.5 = 147.0 g

2. Calculate the actual mass of one molecule of carbon dioxide (CO_2).

Sol. Molecular mass of CO_2 = 44 amu ; 1 amu = 1.66×10^{-24} g

So, the actual mass of CO_2 = $44 \times 1.66 \times 10^{-24}$ = 7.304×10^{-23} g

3. How many molecules of water and oxygen atoms are presen in 0.9 g of water?

Sol. Information shadow :

Mass of water = 0.9 g ; Molar mass of water = 18 g mol^{-1}

Number of molecules of watr and numberof oxygen atoms present in water are to be calculated.

Problem solving strategy :

Number of moles, n = $\dfrac{\text{Mass}}{\text{Molar mass}}$; Number of molecules = $n \times 6.02 \times 10^{23}$

Working it out : $n = \dfrac{0.9}{18} = 0.05$

Number of molecules of water = $0.05 \times 6.02 \times 6.02 \times 10^{23}$ = $3.01 \times 10^{22.}$

As one molecule of water contains one oxygen atom, So, number of oxygen atoms in 3.01×10^{22} molecules of water = 3.01×10^{22}.

4. What is the mass of 3.01×10^{22} molecules of ammonia?

Sol. Gram-molecular mass of ammonia = 17 g ; Numbe of molecules in 17 g (one mole) of NH_3 be *x* g.

So, $\dfrac{3.01 \times 10^{22}}{6.02 \times 10^{23}} = \dfrac{x}{17}$ or $x = \dfrac{17 \times 3.01 \times 10^{22}}{6.02 \times 10^{23}} = 0.85\ g$

5. From 200 mg of CO_2, 10^{21} molecules are removed. How many moles of CO_2 are left?

Sol. Gram-molecules of CO_2 = 44 g

Mass of 10^{21} molecules of CO_2 = $\dfrac{44}{6.02 \times 10^{23}} \times 10^{21}$ = 0.073 g

Mass of CO_2 left = (0.2 –0.073) = 0.127 g ; Number of moles of CO_2 left = $\dfrac{0.127}{44} = 2.88 \times 10^{-3}$

6. How many molecules and atoms of oxygen are present in 5.6 litres of oxygen (O_2) at NTP?

Sol. We know that 22.4 litres of oxygen at NTP contain 6.02×10^{23} molecules of oxygen.

So, 5.6 litres of oxygen at NTP contain $= \dfrac{5.6}{22.4} \times 6.02 \times 10^{23}$ molecules $= 1.505 \times 10^{23}$ molecules.

1 molecule of oxygen contains = 2 atoms of oxygen.

So, 1.505×10^{23} moleculesof oxygen contain $= 2 \times 1.505 \times 10^{23}$ atoms $= 3.01 \times 10^{23}$ atoms.

7. How many electrons are present in 1.6 g of methane?

Sol. Gram-molecular mass of methane (CH_4) = 12 + 4 = 16 g

Number of moles in 1.6 g of methane $= \dfrac{1.6}{16} = 0.1$

Number of molecules of methane in 0.1 mole $= 0.1 \times 6.02 \times 10^{22}$

One molecule of methane has = 6 + 4 = 10 electrons So, 6.02×10^{22} molecules of methane have

$= 10 \times 6.02 \times 10^{22}$ electrons $= 6.02 \times 10^{22}$ electrons.

8. Calculate the number of moles in 25 g of calcium carbonate and number of oxygen atoms.

Sol. Formula mass of calcium carbonate ($CaCO_3$) = 100

No. of moles of $CaCO_3 = \dfrac{\text{Mass (in grams)}}{\text{Formula mass}} = \dfrac{25}{100} = 0.25$ mole

No. of oxygen atoms in one mole of $CaCO_3 = 3 \times 6.02 \times 10^{23}$

No. of oxygen atoms in 0.25 mole of $CaCO_3 = 0.25 \times 3 \times 6.02 \times 10^{23} = 4.515 \times 10^{23}$

9. 6 grams of carbon was completely burnt in oxygen. What would be the volume of CO_2 produced at NTP and how many molecules will be present in that gas?

Sol.

$$C + O_2 \rightarrow CO_2$$

12 g — 22.4 litre at NTP

6 g — 11.2 litre at NTP

22.4 litre litre CO_2 at NTP contains $= 6.02 \times 10^{23}$ molecules

11.2 litre CO_2 at NTP contains $= \dfrac{6.02 \times 10^{23} \times 11.2}{22.4} = 3.01 \times 10^{23}$ molecules

10. The electric charge on the electron is 1602×10^{-19} coulomb. How much charge is present on 0.1 mole of Cu^{2+} ions?

Sol. Charge on one mole of electrons $= 6.02 \times 10^{23} \times 1.602 \times 10^{-19}$ coulomb = 965000 coulomb = 1 faraday

Charge on one mole of Cu^{2+} ions $= 2 \times 96500$ coulomb = 2 faraday

Charge on 0.1 mole of Cu^{2+} ions $= 0.1 \times 2 = 0.2$ faraday

PRACTICE PROBLEMS

1. Calculate the average atomic mass of carbon. Carbon has the following three isotopes with relative abundances and masses are shown below :

Isotope	Relative abundance %	Atomic mass amu
^{12}C	98.892	12
^{13}C	1.108	13.0035
^{14}C	2×10^{-18}	14.00317

2. Calculate the number of gram atoms in 1.4 grams of nitrogen.

3. 16.26 milligrams of sample of an element *x* contains 1.66×10^{20} atoms. What is the atomic mass of the element?

4. The density of a gaseous element is 5 times that of oxygen under identical conditions. If the molecule is triatomic. What will be its atomic mass?
5. How many atoms of oxygen are present in 300 g $CaCO_3$?
6. Which of the following will weigh the most?
   a. 3.01×10^{23} carbon atoms **b.** 20 g copper **c.** 5 gram atom of nitrogen
7. How many nitrogen atoms are present in 160 amu of ammonium nitrate?
8. What volume of CCl_4 having density 1.5 g/cc will contain 10^{25} atoms of chlorine?
9. What is the mass of 1 mole of electrons?
10. What is the molar volume of water at 4°C? Density of water at 4°C is 1 g/cc.
11. Calculate the mass of sodium which contains same number of atoms as are pesent in 4 g of calcium. Atomic masses of calcium and sodium are 40 and 23 respectively.
12. How much time would it take to distribute one Avogadro's number of wheat grains if 10^{10} grains are distributed in one second?
13. How many atoms and molecules of phosphorous are present in 124 g P_4?
14. What is the mass of carbon present in 0.5 mole of $K_4[Fe(CN)_6]$?
15. Calculate the volume occupied by 16 g O_2 at NTP.

TRY IT YOURSELF

1. Calculate the mass of **: i.** 1.2 gram atom of oxygen **ii.** 5.2 gram atom of iodine **iii.** 5.6 gram atom of chlorine.
2. Which was the maximum and minimum mass ?
   i. 25.6 g of oxygen (atomic mass = 16) **ii.** 2.56 gram atom of sodium (atomic mass = 23)
   iii. 0.256 gram molecule of water (molecular mass = 18) **iv.** 0.256 gram atom of iodine (atomic mass = 127)
3. Calculate the number of moles of iron in a sample containing 1.0×10^{22} atoms.
4. How many molecules and atoms of phosphorus are present in 0.1 mole of P_4 molecules?
5. Calculate the number of atoms of hydrogen, oxygen and sulphur in 0.2 mole of sulphuric acid (H_2SO_4).
6. How many moles of gold are present in 49.25 g of gold rod?
7. Calculate the mass of the following :
   i. 1 atom of carbon **ii.** 1 atom of silver **iii.**1 molecule of benzene (C_6H_6) **iv.** 1 molecule of water (H_2O)
8. Which of the following has the maximum mass
   i. 20 g of phosphorus **ii.** 5 moles of water **iii.** 12×10^{24} atoms of hydrogen
9. Calculate the number of molecules and number of atoms present in 11.2 litre of oxygen (O_2) at N.T.P.
10. Calculate the number of molecule present in 34.2 g of cane sugar [$C_{12}H_{22}O_{11}$].
11. Calculate the number of molecules present in : **a.** 1 kg oxygen **b.** 1 dm^3 of hydrogen at N.T.P.
12. Calculate the mass of
    i. 0.1 mole of KNO_3 **ii.** 1×10^{23} molecules of methane **iii.** 112 cm^3 of hydrogen at S.T.P.
13. The atomic mass of a metal M is 56. Calculate the empirical formula of its oxide containing 70% metal.
14. An organic compound was found to contain the following constituents **:**
    C = 33% ; H = 4.7% ; N = 13.2%, Cl = 33.4%. Determine the empirical formula of the compound.
15. An organic compound o analysis gave the following percentage composition :
    C = 57.8%; H = 3.6% and the rest is oxygen. The vapour density of the compound was found to be 83. Find the molecular formula of the compound.
16. How many moles and how many grams of sodium chloride are present in 250 mL of 0.5 M NaCl solution?

STOICHIOMETRY AND REACTIONS IN SOLUTION

- A chemical equation represents a chemical reaction. The equation or a reaction is said to be balanced when it is consistent with the conservation of mass and conservation of electric charge. A balanced chemical equation gives information about relationship between moles and masses of reactants and products in a chemical reaction. The quantitative study of the reactants and the products formed in a reaction is called ***'stoichiometry'.***
- The solution contains two component, solute & solvent. The component which is present in lesser amount is known as solute and the component which is present in large amount is known as solvent.
- The amount of solute present in a given quantity of solvent is expressed in terms of concentration.
- The concentration of solute can be expressed in any of the following ways.

  (i) Mass percent (ii) Mass fraction (iii) Molarity (iv) Molality (v) Normality
- Molarity(M) is number of moles of solute (n) dissolved in a litre of solution $M = \dfrac{n \times 1000}{\text{Volume of solution in cc or ml}}$
- When mass percentage of solute and density of solution is given then molarity is expressed as

  $$M = \frac{\%\text{ age of solute} \times 10 \times \text{density of solution}}{\text{Molecular weight of solute}}$$
- Molality (m) is number of moles of solute (n) dissolved in 1kg of solvent

  $$\text{Molality (m)} = \frac{n \times 1000}{\text{wt of solvent (in gram)}} = \frac{\text{weight of solute} \times 1000}{\text{M. wt of solute} \times \text{wt of solvent (in gram)}}$$

  Molality is independent of temperature.
- Mole fraction s the number of moles of one component (solute/solvent) divided by total number of moles of all the components (solution). $x_A = n_A/n_A + n_B$

  where n_A and n_B are number of moles of two components in solution. Also $x_A + x_B = 1$.
- Relation between mole fraction and molality $X_{solute} = \dfrac{\text{Molality}}{\text{Molality} + 1000/\text{Mol. mass of solvent}}$
- For dilution of concentrated solution number of moles of solute will remain the same therefore, we apply

  $M_1V_1 = M_2V_2$
- When same solution of different concentrations and volume are added and concentration of the resulting mixture has to be found out; then we apply $M_1V_1 + M_2V_2 = M_3(V_1 + V_2)$

  Also, Molarity × Molecular mass of Solute = Normality × Gram equivalent mass of solute
- Milimoles = Molarity × Volume in cc.
- Gram equivalents = Normality × Volume in cc
- Strength = Grams of solute per litre of solution
- $$\boxed{\text{Mass of Substance}} \underset{\xrightarrow[\text{Molar Mass}]{\text{Divided by}}}{\xleftarrow{\text{Multiplied by}}} \boxed{\textbf{Mole}} \underset{\xleftarrow[\text{Avogadro Constant}]{\text{Divided by}}}{\xrightarrow{\text{Multiplied by}}} \boxed{\text{Number of Entities}}$$

  Schematic diagram for interconversion of mass, mole and number of entities.
- For any chemical reaction, the amount of product which may be obtained is limited by the amount of starting material. Even if an excess of one reactant is added, the amount of product formed is determined by the limiting reagent, i.e., the reactant not present in excess.

- Equivalent mass of elements = At. mass/valency
- Eq. wt. of an acid = $\dfrac{\text{Molecular mass}}{\text{Basicity of acid}}$;
- Eq. wt of a base = $\dfrac{\text{Molecular mass}}{\text{Acidity of base}}$
- Equivalent mass for salts = $\dfrac{\text{Formula mass}}{\text{(Valency of cation) (No. of cations)}}$
- Equivalent mass for oxidising agents = $\dfrac{\text{Formula mass}}{\text{No. of electrons gained per molecule}}$
- Equivalent mass for reducing agents = $\dfrac{\text{Formula mass}}{\text{No. of electrons lost per molecule}}$

Empirical and molecular formulae of some compound are given below :

Compound	**Empirical Formula**	**Molecular Formula**	***n***
Hydrogen peroxide	HO	H_2O_2	2
Glucose	CH_2O	$C_6H_{12}O_6$	6
Benzene	CH	C_6H_6	6
Acetic acid	CH_2O	$C_2H_4O_2$	2
Ammonia	NH_3	NH_3	1

Limiting Reactant or Limiting Reagent

Sometimes we find that not all of the reactants are completely consumed in a chemical reaction.

The **limiting reactant** (or limiting reagent) is the reactant that is entirely consumed when a reaction goes to completion. A reactant that is not completely consumed is referred to as an excess reactant. Once one of the reactants is used up, the reaction stops.

It means that moles of the product are always determined by the starting moles of limiting reactant.

How to detect which of the reactants is limiting? It is the basic question which must arise in the mind of the student. A tip for its answer is given below :

Calculate the amount of product that would be produced if each reactant were completely converted to product. The reactant that gives the smallest amount of product is the limiting reactant and will be completely consumed.

The limiting reactant limits the amount of product that can be formed and determines the theoretical yield of the reaction.

Theoretical Yield, Actual Yield and Percentage Yield

The **theoretical yield** of a reaction is the amount of the product produced by the balanced equation when whole of the limiting reagent has reacted. The theoretical yield is, thus, the maximum obtainable yield. In practice, the **actual yield** is always less than the theoretical yield. There are many reasons for this. For instance, many reactions are reversible and so they do not proceed 100 per cent from left to right. Even when reactants are converted totally to products, it may be difficult to recover all the products from the reaction medium. Some reactions are complex in the sense that the products formed may react further among themselves or with the reactants to form still other products. These secondary reactions will reduced the yield of the actual product or products. Therefore, **per cent yield** which describes the proportion of the actual yield to the theoretical yield, is preferred.

I is defined as follows : % yield = $\dfrac{\text{Actual yield}}{\text{Theoretical yield}} \times 100$

1. A solution of ethanol in water is 1.6 molal. How many grams of ethanol are present in 500 g of the solution?

Sol. Molality can be calculated as : $m = \dfrac{w_B \times 1000}{m_B \times w_A}$ ...... (i)

Given m = 1.6, m_B = 46 g mol^{-1} (molar mass of ethanol C_2H_5OH)

Let x g solute w_B is present in solution i.e., $w_B = x \quad \therefore \quad w_A = (500 - x)$

Putting these values in Eq. (i), we get $1.6 = \dfrac{x \times 1000}{46 \times (500 - x)}$; $1.6 \times 500 - 1.6x = \dfrac{1000}{46}x$

$800 = 1.6x + 21.74x = 23.34x$; $x = \dfrac{800}{23.34} = 34.27\,g$. Thus, 34.27 g solute present in the solution.

2. The solubility of $Ba(OH)_2 .8H_2O$ in water at 288 K is 5.6 g per 100 g of water. What is the molality of the hydroxide ions in the saturated solution of barium hydroxide at 288 K? (At. masses of Ba = 137, O = 16, H = 1).

Sol. Molecular mass of $Ba(OH)_2 . 8H_2O$ = 137 + 34 + 144 = 315.

No. of moles in 5.6 g of $Ba(OH)_2 . 8H_2O = \dfrac{5.6}{315} = 0.018$ mole

No. of OH^- moles in 5.6 g of $Ba(OH)_2 . 8H_2O = 2 \times 0.018 = 0.036$ mole

Mass of solvent = 100 g = 0.1 kg ; Molality of OH^- ions $= \dfrac{0.036}{0.1} = 0.36$ m

3. Concentrated aqueous sulphuric acid is 98% H_2SO_4 by mass and has a density of 1.84 g cm^{-3}. What volume of the concentrated acid is required to mole 5 litres of o.5M H_2SO_4?

Sol. $M = \dfrac{x \times d \times 10}{m_B} = \dfrac{9.8 \times 1.84 \times 10}{98} = 18.4$

Before dilution After dilution

$M_1V_1 = M_2V_2$

$18.4 \times V_1 = 0.5 \times 5000$; $V_1 = 135.86$ mL

4. Calculate the percentage composition of calcium nitrate.

Sol. The formula of calcium nitrate is $Ca(NO_3)_2$.

Thus, the formula mass or molecular mass = At. mass of Ca + 2 × At. mass of N + 6 × At. mass of O

$= 40 + (2 \times 14) + (6 \times 16) = 164$; % of Ca $= \dfrac{40}{164} \times 100 = 24$; % of N $= \dfrac{28}{164} \times 100 = 17$;

% of oxygen = 100 –(24 + 17) = (100 –41) = 59

5. Calculate the empirical formula of a compound that contains 26.6% potassium, 35.4% chromium and 38.1% oxygen. [Given : K = 39.1; Cr = 52; O = 16].

Sol.

Element	**%**	**Atomic mass**	**Relative no. of atoms**	**Simplest ratio**	**Simplest whole no. ratio**
Potassium	26.6	39.1	$\dfrac{26.6}{39.1} = 0.68$	$\dfrac{0.68}{0.68} = 1$	1 × 2 = 2
Chromium	35.4	52.0	$\dfrac{35.4}{52} = 0.68$	$\dfrac{0.68}{0.68} = 1$	1 × 2 = 2
Oxygen	38.1	16.0	$\dfrac{38.1}{16} = 2.38$	$\dfrac{2.38}{0.68} = 3.5$	3.5 × 2 = 7

Therefore, empirical formula is $K_2Cr_2O_7$.

6. An organic compound on analysis gave the following percentage composition : C = 57.8%, H = 3.6% and the rest is oxygen. The vapour density of the compound was found to be 83. Find out the molecular formula of the compound.

Sol.

Element	**%**	**Atomic mass**	**Relative no. of atoms**	**Simplest ratio**	**Simplest whole no. ratio**
Carbon	57.8	12	4.82	$\frac{4.82}{2.41} = 2$	4
Hydrogen	3.6	1	3.6	$\frac{3.6}{2.41} = 1.49$	3
Oxygen	38.6	16	2.41	$\frac{2.41}{2.41} = 1$	2

Empirical formula = $C_4H_3O_2$; Empirical mass = $4 \times 12 + 3 \times 1 + 2 \times 16 = 83$;

Molecular mass = 2V.D = $2 \times 83 = 166$; $n = \frac{\text{Molecular mass}}{\text{Empirical mass}} = \frac{166}{83} = 2$

Molecular formula = n × Empirical formula = $2 \times C_4H_3O_2 = C_8H_6O_4$

7. Ethyl alcohol is prepared industrially by the reaction of ethylene (C_2H_4) with water.

$$\underset{\text{Ethylene}}{C_2H_4}(g) + H_2O(l) \longrightarrow CH_3CH_2OH(l)$$

What is the percentage yield of the reaction if 4.6 g of ethylene gives 4.7 g of ethyl alcohol?

Sol. $$\underset{\substack{\text{Ethylene}\\ 1\,\text{mol}\,(28\,\text{g})}}{C_2H_4}\ (g) + H_2O(l) \longrightarrow \underset{\substack{\text{Ethyl alcohol}\\ 1\,\text{mol}\,(46\text{g})}}{CH_3CH_2OH(l)}$$

Mass of ethyl alcohol formed by 4.6 g ethylene $= \frac{46}{28} \times 4.6 = 7.557\ g$

Percentage yield $= \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100 = \frac{4.7}{7.557} \times 100 = 62.19\ \%$

8. Hydrogen chloride (HCl) an oxidation gives water and chlorine. How many litres of chlorine at STP can be obtained starting with 36.50 g HCl?

Sol. Oxidation of HCl takes place according to the following equation :

$$\underset{4\,\text{mol}}{4HCl} + O_2 \longrightarrow \underset{2\,\text{mol}}{2Cl_2} + 2H_2O \quad ; \quad \text{Moles of HCl} = \frac{\text{Mass}}{\text{Molecular mass}} = \frac{36.5}{36.5} = 1\,\text{mol.}$$

$\because$ 4 moles HCl give 2 moles Cl_2 ; $\because$ 1 mole will give 2/4 moles Cl_2 = 0.5 mole Cl_2

Volume of Cl_2 at STP = 22.4×0.5 = 11.2 litre.

9. Reaction : $2Br^-(aq) + Cl_2(aq) \longrightarrow 2Cl^-(aq) + Br_2(aq)$ is used for commercial preparation of bromine from its salts. Suppose we have 50 mL of 0.06 M NaBr. What volume of 0.050 M solution of Cl_2 is needed to react completely with the Br^-?

Sol. The given reaction is : $\underset{n_1 = 2}{2Br^-}(aq) + \underset{n_2 = 1}{Cl_2}(aq) \longrightarrow 2Cl^-(aq) + Br_2(aq)$

$$\frac{M_1V_1}{n_1} = \frac{0.05 \times V_2}{n_2} \quad \therefore \quad \frac{0.06 \times 50}{2} = \frac{0.05 \times V_2}{1} \quad ; \quad V_2 = 30\ \text{mL}$$

10. When C_4H_{10} is burned in excess oxygen, the following reaction occurs :

$2C_4H_{10}(g) + 13O_2(g) \longrightarrow 8CO_2(g) + 10H_2O(g)$

What volume of oxygen at NTP is required to burn 36 g of C_4H_{10}?

Sol. The given reaction shows that : 2 mol C_4H_{10} requires 13 mol O_2 i.e., 2×58 g C_4H_{10} requires 13×22.4L O_2 at STP

$\therefore$ 36 g C_4H_{10} will require $\left[\dfrac{13\times 22.4\times 36}{2\times 58}\right]$L O_2 at NTP

Thus, required volume of O_2 = $\dfrac{13\times 22.4\times 36}{2\times 58}$ L O_2 = 90.37 L at STP.

PRACTICE PROBLEMS

1. Calculate the molality of a solution of ethanol in water in which the mole fraction of ethanol is 0.04.

2. A solution is obtained by mixing 300 g of 25% solution and 400 g of 40% solution by mass. Calculate mass percentage of resulting solution.

3. Calculate the percentage composition of water of crytallisation in the pure sample of green vitriol $FeSO_4 \cdot 7H_2O$.

4. Atomic mass of iron is 56; its oxide contains 70% iron. Calculate the empirical formula of oxide.

5. A compound on analysis gave following percentage composition : C = 57.8%, H = 3.6% and the remainder is the oxygen. The vapour density of the compound is 83. Find the molecular formula of the compound.

6. Calculate the mass of 60% H_2SO_4 required to decompose 50g chalk $CaCO_3$.

7. 1 g of Mg is burnt in a closed vessel which contains 0.5 g of O_2.

a. Which reactant is left in excess? **b.** Find the mass of excess reactant

8. Calculate the nu7mber of grams of $MgCl_2$ that could be obtained from 17g HCl in following reaction :

$MgO + 2HCl \longrightarrow MgCl_2 + H_2O$.

9. If 20 g of $CaCO_3$ is treated with 20 g of HCl, how many grams of CO_2 can be generated according to the following equation?

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

10. How many grams of oxygen are required to burn completely 570 g of octane?

TRY IT YOURSELF

1. Calculate the weight of iron which will be converted into its oxide (Fe_3O_4) by the action of 18 g of steam on it.

2. 1.0 g of Mg is burnt in a closed vessel which contains 0.5 g of O_2. Which is the limiting reactant. What is the amount of MgO formed in the reaction?

3. A solution is prepared by dissolving 18.25 g of NaOH in 200 mL of it. Calculate the molarity of the solution.

4. A bottle of commercial sulphuric acid (density = 1.787 gm/mL) is labelled 86 percent by weight. What is the molarity of the solution? What volume of the acid is required to make it 1 litre of 0.2 M H_2SO_4?

5. What will be the normality of the solution obtained by adding 200 mL of pure water to 100 mL of 0.2 N HCl solution?

5. What volume of 10 M HCl and 3 M HCl should be mixed to get 1 L of 6 M HCl solution?

6. Calculate the molality of a solution containing 20.7 g of potassium carbonate in 500 mL of solution (assume density of solution = 1 g mL^{-1})

TEXT BOOK QUESTIONS

1. Calculate the molecular mass of the following:

(i) H_2O (ii) CO_2 (iii) CH_4

2. Calculate the mass per cent of different elements present in sodium sulphate (Na_2SO_4).

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

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

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

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

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

6. Calculate the concentration of nitric acid in moles per litre in a sample which has a density, 1.41 g mL^{-1} and the mass per cent of nitric acid in it being 69%.

7. How much copper can be obtained from 100g of copper sulphate ($CuSO_4$)?

8. Determine the molecular formula of an oxide of iron in which the mass per cent of iron and oxygen are 69.9 and 30.1 respectively.

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

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

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

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

(iii) Number of molecules of ethane.

11. What is the concentration of sugar ($C_{12}H_{22}O_{11}$) in mol L^{-1} if its 20 g are dissolved in enough water to make a final volume up to 2L?

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

13. Pressure is determined as force per unit area of the surface. The SI unit of pressure, pascal is as shown below :

1Pa = 1N m^{-2}

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

14. What is the SI unit of mass? How is it defined?

15. Match the following prefixes with their multiples:

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

16. What do you mean by significant figures?

17. A sample of drinking water was found to be severely contaminated with chloroform, $CHCl_3$, supposed to be carcinogenic in nature. The level of contamination was 15 ppm (by mass).

(i) Express this in percent by mass. (ii) Determine the molality of chloroform in the water sample.

18. Express the following in the scientific notation: (i) 0.0048 (ii) 234,000 (iii) 8008 (iv) 500.0 (v) 6.0012

19. How many significant figures are present in the following?

(i) 0.0025 (ii) 208 (iii) 5005 (iv) 126,000 (v) 500.0 (vi) 2,0034

20. Round up the following upto three significant figures: (i) 34.216 (ii) 10.4107 (iii) 0.04597 (iv) 2808

21. The following data are obtained when dinitrogen and dioxygen react together to form different compounds:

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

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

b. Fill in the blanks in the following conversions:

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

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

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

23. In a reaction $A + B_2 \longrightarrow AB_2$. Identify the limiting reagent, if any, in the following reaction mixtures.

(i) 300 atoms of A + 200 molecules of B (ii) 2 mol A + 3 mol B

(iii) 100 atoms of A + 100 molecules of B (iv) 5 mol A + 2.5 mol B (v) 2.5 mol A + 5 mol B

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

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

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

(ii) Will any of the two reactants remain unreacted? (iii) If yes, which one and what would be its mass?

25. How are 0.50 mol Na_2CO_3 and 0.50 M Na_2CO_3 different?

26. If ten volumes of dihydrogen gas reacts with five volumes of dioxygen gas, how many volumes of water vapour would be produced?

27. Convert the following into basic units:

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

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

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

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

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

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

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

32. Use the data given in the following table to calculate the molar mass of naturally occuring argon isotopes:

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

33. Calculate the number of atoms in each of the following (i) 52 moles of Ar (ii) 52 u of He (iii) 52 g of He.

34. A welding fuel gas contains carbon and hydrogen only. Burning a small sample of it in oxygen gives 3.38 g carbon dioxide, 0.690 g of water and no other products. A volume of 10.0 L (measured at STP) of this welding gas is found to weigh 11.6 g. Calculate (i) empirical formula, (ii) molar mass of the gas, and (iii) molecular formula.

35. Calcium carbonate reacts with aqueous HCl to give $CaCl_2$ and CO_2 according to the reaction.

$CaCO_3(s) + 2HCl\,(aq) \longrightarrow CalCl_2\,(aq) + CO_2(g) + H_2O(\ell)$

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

36. Chlorine is prepared in the laboratory by treating manganese dioxide (MnO_2) with aqueous hydrochloric acid according to the reaction $4\,HCl(aq) + MnO_2(s) \longrightarrow 2H_2O(\ell) + MnCl_2(aq) + C_2(g)$. How many grams of HCl react with 5.0 g of manganese dioxide?

EXERCISE

MULTIPLE CHOICE QUESTIONS

TOPIC–1 : GENERAL INTRODUCTION AND NATURE OF MATTER

1. Two students performed the same experiment separately and each one of them recorded two readings of mass which are given below. Correct reading of mass is 3.0 g. On the basis of given data, mark the correct option out of the following statements.

Students	**Reading**	
	(i)	**(ii)**
A	3.01	2.99
B	3.05	2.95

a. Results of both the students are neither accurate nor precise.

b. Results of student A are both precise and accurate.

c. Results of student B are neither precise nor accurate. d. Results of student B are both precise and accurate.

2. A measured temperature on Fahrenheit scale is 200 °F. What will this reading be on Celsius scale?

a. 40°C b. 94°C c. 93.3°C d. 30° C

3. Which of the following is an example of a heterogeneous substance?

a. Bottled water b. Table salt c. Piece of copper d. Candle

4. Which of the following statements about a compound is incorrect?

a. A molecule of a compound has atoms of different elements.

b. A compound cannot be separated into its constituent elements by physical methods of separation.

c. A compound retains the physical properties of its constituent elements.

d. The ratio of atoms of different elements in a compound is fixed.

5. The prefix 10^{18} is

a. giga b. kilo c. exa d. nano

6. Given the numbers: 161 cm, 0.161 cm, 0.0161 cm. The number of significant figures for the three numbers are

a. 3, 4 and 5 respectively b. 3, 3 and 4 respectively c. 3, 3 and 3 respectively d. 3, 4 and 4 respectively

7. If the density of a solution is 3.12 g mL^{-1}, the mass of 1.5 mL solution in significant figures is_________.

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

8. In which of the following number all zeroes are significant? a. 0.0005 b. 0.0500 c. 50.000 d. 0.0050

TOPIC–2 : LAWS OF CHEMICAL COMBINATION

9. Which of the following statements is correct about the reaction given below? $4Fe(s) + 3O_2(g) \longrightarrow 2Fe_2O_3(g)$

a. Total mass of iron and oxygen in reactants = total mass of iron and oxygen in product therefore, it follows law of conservation of mass.

b. Total mass of reactants = total mass of product; therefore, law of multiple proportions is followed.

c. Amount of Fe_2O_3 can be increased by reducing the amount of any one of the reactants (iron or oxygen).

d. Amount of Fe_2O_3 produced will decrease if the amount of any one of the reactants (iron or oxygen) is taken in excess.

10. Two samples of lead oxide were separately reduced to metallic lead by heating in a current of hydrogen. The weight of lead from one oxide was half the weight of lead obtained from the other oxide. The data illustrates –

a. law of reciprocal proportions b. law of constant proportions

c. law of multiple proportions d. law of equivalent proportions

11. In compound A, 1.00g of nitrogen unites with 0.57 g of oxygen. In compound B, 2.00g of nitrogen combines with 2.24g of oxygen. In compound C, 3.00g of nitrogen combines with 5.11g of oxygen. These results obey the following law

a. law of constant proportion   b. law of multiple proportion
c. law of reciprocal proportion   d. Dalton's law of partial pressure

12. One mole of a gas occupies a volume of 22.4 L. This is derived from
a. Berzelius' hypothesis b. Gay-Lussac's law c. Avogadro's law d. Dalton's law

TOPIC–3 : DIFFERENT MASSES, MOLE CONCEPTS AND FORMULAE OF COMPOUNDS

13. Molecular mass is defined as the
a. mass of one atom compared with the mass of one molecule
b. mass of one atom compared with the mass of one atom of hydrogen
c. mass of one molecule of any substance compared with the mass of one atom of C-12   d. None of the above

14. The empirical formula and molecular mass of a compound are CH_2O and 180 g respectively. What will be the molecular formula of the compound?
a. $C_9H_{18}O_9$ b. CH_2O c. $C_6H_{12}O_6$ d. $C_2H_4O_2$

15. What is the mass percent of carbon in carbon dioxide?
a. 0.034% b. 27.27% c. 3.4% d. 28.7%

TOPIC–4 : STOICHIOMETRY AND REACTIONS IN SOLUTION

16. Two containers P and Q of equal volume (1 litre each) contain 6 g of O_2 and SO_2 respectively at 300 K and 1 atmosphere, then
a. Number of molecules in P is less than that in Q   b. Number of molecules in P and Q is same
c. Number of molecules in Q is less than that in P   d. Either a or b

17. 50 ml 10 N H_2SO_4, 25 ml 12 N HCl and 40 ml 5 N HNO_3 were mixed together and the volume of the mixture was made 1000 ml by adding water. The normality of the resultant solution will be
a. 2 N b. 1 N c. 3 N d. 4 N

18. In a chemical reaction : $K_2Cr_2O_7 + xH_2SO_4 + ySO_2 \longrightarrow K_2SO_4 + zCr_2(SO_4)_3 + H_2O$ the values of x, y, z are
a. 4, 1, 4 b. 1, 3, 1 c. 3, 2, 3 d. 2, 1, 2

19. 25.4 g of I_2 and 14.2 g of Cl_2 are made to react completely to yield a mixture of ICl and ICl_3. Calculate mole of ICl and ICl_3 formed
a. 0.1, 0.1 b. 0.2, 0.2 c. 0.1, 0.2 d. 0.2, 0.1

20. The Statue of Liberty is made of 2.0×10^5 lbs of copper sheets bolted to a framework. (11b = 454 g) How many atoms of copper are on the statue?   (Atomic weight: Cu = 63.5).
a. 2.1×10^{27} b. 8.6×10^{29} c. 4.3×10^{26} d. 8.6×10^{26}

VERY SHORT ANSWER QUESTIONS

TOPIC–1 : GENERAL INTRODUCTION AND NATURE OF MATTER

1. How many significant figures are present in 6.023×10^{23}?
2. Write 0.0005729 in scientific notation.
3. Q is an impure substance, is it an element, compound or a mixture?
4. Express the following results to the proper number of significant figures. $\dfrac{(1.38\times10^{-4})(0.8)}{3.9}$
5. Why air is not always regarded as homogeneous mixture?
6. Vanadium metal is added to steel to impart strength. The density of vanadium is 5.96 g/cm^3. Express this in SI unit.
7. What is the difference between 2H and H_2?
8. What are basic properties used to identify a substance?
9. What is the S.I. unit of energy?
10. State the number of significant figures in each of the following numbers. (i) 0.05031 (ii) 2.563×10^{10}

TOPIC–2 : LAWS OF CHEMICAL COMBINATION

11. Is the law of constant composition true for all types of compounds? Explain why or why not.

12. 45.4L of dinitrogen reacted with 22.7 L of dioxygen and 45.4 L of nitrous oxide was formed. The reaction is given below: $2N_2(g) + O_2(g) \longrightarrow 2N_2O(g)$

Which law is being obeyed in this experiment? Write the statement of the law.

13. Nitrogen and Oxygen form N_2O, NO, NO_2, N_2O_4, N_2O_5. Which Law of chemical combination is followed?

14. Which element is used as a standard for comparing atomic and molecular masses?

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

16. Dinitrogen combines with dihydrogen to form ammonia according to the following reaction.

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

What is the ratio of their volumes under similar conditions of temperature and pressure?

TOPIC–3 : DIFFERENT MASSES, MOLE CONCEPTS AND FORMULAE OF COMPOUNDS

17. Calculate the mass of one atom of carbon?

18. Calculate the total number of electrons present in 1.4 g of nitrogen gas?

19. A colourless liquid used in rocket engines, whose empirical formula is NO_2, has a molar mass of 92. What is its molecular formula?

20. What is the mass of 2 atoms of oxygen?

21. Write the formulae and names of three compounds containing same percentage composition of C, H and O.

22. Why are the atomic masses of most of the elements fractional?

23. How many grams of silver are in 0.263 mol of Ag?

24. Calculate the mass of 6.022 × 1022 atoms of He.

25. Describe the difference between the mass of a mole of oxygen atoms (O) and the mass of a mole of oxygen molecules (O_2).

26. Boron occurs in nature in the form of two isotopes $^{11}_{5}B$ and $^{10}_{5}B$ in ratio of 81% and 19% respectively. Calculate its average atomic mass.

27. 0.5 mole each of H_2S and SO_2 mixed together in a reaction flask, react according to equation:

$2H_2S + SO_2 \rightarrow 2H_2O + 3S$. Calculate the number of moles of 'S' formed.

28. Calculate the mass of ferric oxide that will be obtained by complete oxidation of 2 g of Fe.

[Atomic weights of Fe = 56 u, O = 16 u]

TOPIC–4 : STOICHIOMETRY AND REACTIONS IN SOLUTION

29. Reaction, $2Br^-(aq) + Cl_2 (aq) \longrightarrow 2Cl^-(aq) + Br^2(aq)$, is used for commercial preparation of bromine from its salts. Suppose we have 50.0 ml of a 0.060M solution of NaBr. What volume of a 0.050 M solution of Cl_2 is needed to react completely with Br^-?

30. Volume of a solution changes with change in temperature, then, will the molality of the solution be affected by temperature? Give reason for your answer.

31. In what smallest whole-number ratio must N and O atoms combine to make dinitrogen to tetroxide N_2O_4? What is the mole ratio of the elements in this compound?

32. Alkaline solution of $KMnO_4$ reacts as follows: $2KMnO_4 + 2KOH \longrightarrow 2K_2MnO_4 + H_2O + [O]$ Calculate the equivalent weight of $KMnO_4$ in basic medium.

33. Balance the equation $C_3H_8(g) + O_2 (g) \longrightarrow CO_2 (g) + H_2O(l)$

34. Balance the equation $P_4(s) + O_2(g) \longrightarrow P_4O_{10}(s)$

SHORT ANSWER QUESTIONS

TOPIC–1 : GENERAL INTRODUCTION AND NATURE OF MATTER

1. Express the following numbers upto three significant figures:

(i) 5.607883 (ii) 32.394800 (iii) six thousand (iv) 0.007838

2. What is derived unit? Derive the unit for density in SI unit?

3. The population of India based on 1981 census figure was 684 million. Express the results in scientific notation and calculate the number of significant figures.

4. a. Classify the following as pure substances or mixtures.

b. Separate the pure substance into elements and compounds and divide the mixtures into homogeneous and heterogeneous categories:

(i) graphite (ii) milk (iii) air (iv) diamond (v) petrol (vi) tap water (vii) distilled water (viii) oxygen (ix) 22 carat gold (x) steel (xi) iron (xii) iodized table salt (iii) wood (xiv) cloud.

5. Give at least four points of difference between a compound and a mixture.

6. Explain how compounds differ from elements?

7. Explain how mixture differs from pure substance?

8. a. How many significant figures are there in 1.00×10^6?

b. One mole of sugar contains .... oxygen atoms.

c. Give an example of molecule in which the empirical formula is CH_2O and the ratio of molecular formula weight and empirical formula weight is 6.

9. What are various types of mixtures? Give at least one example in each case?

TOPIC–2 : LAWS OF CHEMICAL COMBINATION

10. Is the Law of constant composition true for all types of compounds? If not so, then why?

11. a. When 4.2 g of $NaHCO_3$ (sodium hydrogen carbonate) is added to a solution of CH_3COOH (acetic acid) weighing 10.0 g, it is observed that 2.2 g CO_2 is released into atmosphere. The residue is found to weigh 12.0 g. Show that three observations are in agreement with the law of conservation mass.

b. If 6.3 g of $NaHCO_3$ are added to 15.0 g CH_3COOH solution, the residue is found to weigh 18.0 g. What is the mass of CO_2 released in the reaction?

12. If two elements can combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element, are in whole number ratio.

a. Is this statement true? b. If yes, according to which law? c. Give one example related to this law.

TOPIC–3 : DIFFERENT MASSES, MOLE CONCEPTS AND FORMULAE OF COMPOUNDS

13. How many molecules of water of hydration are present in 39.2 mg of Mohr's salt? [$FeSO_4 \,.\, (NH_4)_2SO_4 .\, 6H_2O$]

14. One second is required to count two wheat grain, then calculate the time (in years) required to count one mole of wheat grains.

15. Naphthalene contains 93.7% carbon and 6.29% hydrogen, if its molecular mass is 128 g/mol. Calculate its molecular formula.

16. How many moles of hydrogen, phosphorous and oxygen are present in 0.213 moles of H_3PO_4 (Phosphoric acid)?

17. When 36g of H_2O reacted with Fe it gave 90.0g of Fe_3O_4. What is the % yield of the reaction?

$$3\,Fe + 4H_2O \longrightarrow Fe_3O_4 + 4H_2$$

18. A 0.1g sample of a compound when burnt completely in oxygen produced 0.191g of CO_2 and 0.1172g of H_2O. What is the empirical formula of the compound?

19. Two bulbs B_1 and B_2 of equal capacity contain 10g oxygen (O_2) and ozone (O_3) respectively. Which bulb will have greater number of O-atoms and which will have greater number of molecules?

20. Calculate the number of atoms of the constituent elements is 53 g of Na_2CO_3.

21. 4 g of copper chloride on analysis was founded to contain 1.890 g of copper (Cu) and 2.110 g of chlorine (Cl). What is the empirical formula of copper chloride? [Atomic mass of Cu = 63.5 u, Cl = 35.5 u]

22. The cost of table salt (NaCl) and table sugar ($C_{12}H_{22}O_{11}$) are Rs. 2 per kg and Rs. 6 per kg respectively calculate their cost per mole.

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

24. What do you mean by the term 'Formality'? To what type of compounds it is applied?

25. Calculate the number of moles in each of the following.

(i) 392 g of sulphuric acid (ii) 44.8 litres of sulphur dioxide at N.T.P. (iii) 8g of calcium

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

27. $Fe_2(SO_4)_3$ is empirical formula of a crystalline compound of iron. It is used in water and sewage treatment to acid in the removal of suspended impurities. Calculate the mass percentage of iron, sulphur and oxygen in this compound.

28. Calculate the number of electrons present in 21g of nitride ions.

TOPIC–4 : STOICHIOMETRY AND REACTIONS IN SOLUTION

29. Zinc and hydrochloric acid react according to the equation : $Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$

If 0.30 mole of Zn are added to hydrochloric acid containing 0.52 mole of HCl. Which of the two reactant is limiting reagent and how many moles of H_2 are produced?

30. Calculate the molality of 1M NaOH solution which has a density of 1.2 g/mL?

31. Calculate the weight of CaO obtained by heating 200g of 95% pure limestone, $CaCO_3$?

32. A glass of juice contains 9 gm of glucose ($C_6H_{12}O_6$). How many atoms of each element (C, H and O) are there in the juice.

33. From 6.02×10^{22} molecules of N_2 present in a container, 700 mg of N_2 are removed. What is the amount of H_2 in grams required to convert the remaining N_2 into NH_3?

34. $BaCl_2$ solution has a density of 1.279 g/mL and the percent of solute is 26%. Calculate the molality of the solution? (Mol. Weight of $BaCl_2$ = 208)

35. Calculate the molality of a sulphuric acid solution in which mole fraction of water is 0.85.

36. What information is conveyed by the following chemical equation? $Mg + 2HCl \rightarrow MgCl_2 + H_2$.

37. If 20 g of $CaCO_3$ is treated with 20 gram of HCl, how many grams of CO_2 can be generated according to given equation: $CaCO_3(s) + 2HCl\ (aq) \rightarrow CaCal_2(aq) + H_2O(l) + CO_2(g)$

38. a. What is limiting reactant?

b. Oxygen is prepared by catalytic decomposition of potassium chlorate ($KClO_3$). Decomposition of potassium chlorate gives potassium chloride (KCl) and oxygen (O_2). If 2.4 mole of oxygen is needed for an experiment, how many grams of potassium chlorate must be decomposed? (At mass K = 39, Cl = 35.5, O=16)

39. Calculate the moles of NaOH required to neutralize the solution produced by dissolving 1.1 g P_4O_6 in water. Use the following reactions:

$P_4O_6 + 6H_2O \rightarrow 4H_3PO_3$; $2NaOH + H_3PO_3 \rightarrow Na_2HPO_3 + 2H_2O$ (Atomic mass/g mol^{-1}; P = 31, O = 16

40. Calculate the weight of FeO formed from 2g of VO and 5.75 g of Fe_2O_3. Also report the limiting reagent.

$2VO + 3Fe_2O_3 \rightarrow 6FeO + V_2O_5$ (Atomic mass V = 51.4, O = 16, Fe = 55.9 g)

41. The reactant which is entirely consumed in reaction is known as limiting reagent. In the reaction $2A + 4B \rightarrow 3C + 4D$, when 5 moles of A react with 6 moles of B, then

(i) which is the limiting reagent (ii) calculate the amount of C formed?

42. What volume of 6M HCl and 2M HCl should be mixed to get two litres of 3M HCl?

LONG ANSWER QUESTIONS

1. (i) 0.9031g of mixture of NaCl and KCl on treatment with conc. H_2SO_4 yields 1.0784g of a mixture of Na_2SO_4 and K_2SO_4. Calculate the % composition of the mixture.

(ii) Calculate the percentage of the naturally occurring isotopes ^{35}Cl and ^{37}Cl that accounts for the mass of chloride taken as 35.45.

2. (i) Concentrated aqueous sulphuric acid is 98% H_2SO_4 by mass and has a density of 1.84 g mL^{-1}. What volume of the concentrated acid is required to make 5.0L of 0.50 M H_2SO_4 solution? (Mol. weight of sulphuric acid = 98)

(ii) You are given a solution of 14.8M NH_3. How many milliliters of this solution do you require to give 100 ml of 1MNH_3? How much of water will you add?

3. (i) Copper oxide was prepared by the following methods:

a. IN one case, 1.75 g of the metal were dissolved in nitric acid and igniting the residual copper nitrate yielded 2.19 g of copper oxide.

b. In the second case, 1.14 g of metal dissolved in nitric acid were precipitated as copper hydroxide by adding caustic alkali solution. The precipitated copper hydroxide after washing, drying and heating yielded 1.43g of copper oxide.

c. In the third case, 1.45 g of copper when strongly heated in a current of air yielded 1.83 g of copper oxide.

Show that the given data illustrate the law of constant composition.

(ii) Elements A and B form two different compounds. In first case 0.52 grams of A combines with 0.72 grams of B and in second case 0.15 grams of A combines with 0.65 grams of B. Show that these data illustrate the Law of multiple proportion.

4. (i) Calculate the volume at STP occupied by a. 14 g of nitrogen, b. 1.5 moles of carbon dioxide of c. 10^{21} molecules of oxygen.

(ii) Ammonia contains 82.35% of nitrogen and 17.65% of hydrogen. Water contains 88.90% of oxygen and 11.10% of hydrogen. Nitrogen trioxide contains 63.15% of oxygen and 36.85% of nitrogen. Show that these data illustrate the law of reciprocal proportions.

5. What are the limitations of a chemical equation? How could these limitations be removed?

6. A compound on analysis gave the following percentage composition:

Na = 14.31%, S = 9.97%, H = 6.22% O = 69.50%

Calculate the molecular formula of the compound on the assumption that all the hydrogen in the compound is present in combination with oxygen as water of crystallization. Molecular mass of the compound is 322.

PART V SKILL ANALYSER

Time: 30 min. **Max. Marks :** 15

Directions: (i) Attempt at questions (ii) Question 1 to 3 carry 1 mark each. (iii) Question 4 and 5 carry 2 marks each. (iv) Question 6 carry 3 marks (v) Question 7 carry 5 marks

1. What are the SI units of molarity?

2. Which is more concentrated 1 molal or 1 molar?

3. A commercially available sample of H_2SO_4 is 15% H_2SO_4 by weight (density = 1.10g/ml). Calculate molarity of the solution.

4. You are given one litre of 0.15 M HCl and one litre of 0.40 M HCl. What is the maximum volume of 0.25 M HCl? Which you can make from these solutions without adding any water?

5. Find out the volume of Cl_2 at STP produced by action of 100 cc of 0.2M HCl on excess of MnO_2.

6. 500 cc of 0.250 M Na_2SO_4 solution added to an aqueous solution of 15.00 g of $BaCl_2$. How many moles and how many grams of $BaSO_4$ are formed? (Mol. wt. Of $BaSO_4$ = 233, and $BaCl_2$ = 208)

7. (i) Two oxides of metal contain 27.6% and 30% of oxygen respectively. If the formula of first oxide is M_3O_4, find that of second.

(ii) Gastric juice contains about 3.0 g of HCl per liter. If a person produces about 2.5 liter of gastric juice per day, how many antacid tablets each containing 400 mg of $Al(OH)_3$ are needed to neutralise all the HCl produced in one day? Atomic mass of Cl = 35.5; Al = 27; O = 16; H = 1.